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Chemistry Teacher's Edition

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CHAPTER 1 The Science of Chemistry TE

Chapter Outline

Unit 1 Introduction to Chemistry

Outline

This unit, *Introduction to Chemistry*, includes two chapters that introduce students to the Science of Chemistry.

- **Chapter 1 Introduction to Chemistry**
- **Chapter 2 Measurement**

Overview

The Science of Chemistry

This chapter details the scientific method while the core of the chapter gives a brief history of chemistry.

Measurement

This chapter covers qualitative versus quantitation observations, measurement, the mathematics of measurement and formulas, and data handling techniques.

The Science of Chemistry

Outline

The Introduction to Chemistry chapter consists of two lessons that detail the scientific method while the core of the chapter gives a brief history of chemistry.

- **Lesson 1.1 What is Chemistry?**
- **Lesson 1.2 The Process of Science**

Overview

In these lessons, students will explore:

- The use of the scientific method.
- The definition and history of chemistry, the law of conservation of mass, and the use of scientific models.
- The role of a chemist as a scientist who studies the properties of matter.
- The concept map below provides a visual representation of how the chapter concepts are related.

Science Background Information

This background information is provided for teachers who are just beginning to teach in this subject area.

What is Mass-Energy Equivalence?

Albert Einstein is best known for his theories of relativity. There are two parts to the theory. The first part is the special theory of relativity, which was proposed in 1905. The second is the general theory of relativity, which was proposed in 1915. Einstein's special theory of relativity describes the motion of particles moving close to the speed of light. Mass-energy equivalence is a consequence of the special theory of relativity. Mass-energy equivalence is the concept that a measured quantity of energy is equivalent to a measured quantity of mass. The formula $E = mc^2$ expresses the connection between mass and energy. Here E represents energy, m represents mass, and c represents the speed of light in a vacuum. Because the speed of light is a very large number $(299,792,458\ m/s)$ and it is squared, the equation shows that very small amounts of mass can be converted into very large amounts of energy and vice versa.

Pacing the Lessons

Use the table below as a guide for the time required to teach the lessons of *The Science of Chemistry*.

TABLE 1.1: Class Periods per Lesson

Lesson	Number of 60 Minute Class Periods
1.1 The Scientific Method	2.0
1.2 Chemistry in History	0.5
1.3 Chemistry is a Science of Materials	2.0
1.4 Matter	1.0
1.5 Energy	1.5

Managing Materials

The following materials are needed to teach the strategies and activities described in the Teachers Edition of the Flexbook for *The Science of Chemistry.*

TABLE 1.2: The Science of Chemistry Materials List

Lesson	Strategy or Activity	Materials Needed
1.2 Chemistry in History	Exploration Activity	vinegar, baking soda, soda bottle, balloon
1.3 Chemistry is a Science of Materials	Exploration Activity	lighter, birthday candle
1.5 Energy	Exploration Activity	glycerin, beaker, metal spoon

Multimedia Resources

You may find these additional web-based resources helpful when teaching *The Science of Chemistry.*

- A list of forms of energy. http://web.singnet.com.sp/ stepchin/Forms.htm

Possible Misconceptions

Identify: Students may confuse theories and hypotheses. This misconception may arise because of the everyday use of the word theory as in, "it's just a theory" or "tell us your theory." Also, some students may relate these terms in a hierarchical manner in that they may think that hypotheses become theories, which in turn become scientific laws. It is important that student are able to correctly define the terms: "hypothesis," "theory" and "law," as well have a clear understanding of the relationships among them.

Clarify: A hypothesis is a proposal intended to explain a set of observations. Not all hypotheses become theories. A theory is a hypothesis that has been supported with repeated testing. A law is a relationship that exists between specific observations. In other words, a law is a relationship that always applies under a given set of conditions.

Promote Understanding: Have students use a dictionary to define these three terms. Explain to students that there is no, "hierarchy of terms." In other words, a theory is not better than a hypothesis, and a law is not better than a theory. Point out that hypotheses, laws and theories each have their place in science. On the board, draw a Venn diagram to illustrate the relationship between a scientific theory and a scientific law. Label the circle on the left, "scientific theory." Label the circle on the right, "scientific law." Have students define each term in the appropriate circle. In the section where the two circles overlap, have students come up with some similarities between a scientific theory and a scientific law.

Discuss: At the end of the lesson ask, "What are some similarities between a scientific theory and a scientific law?" (Both are based on observation and experimentation.)

Ask: What are some differences between a scientific theory and a scientific law? (A theory is more of an explanation whereas a law is just a statement or description of a relationship.)

Making the FlexBook Flexible

An important advantage of the FlexBook is the ability it gives you, the teacher, to select the chapters and lessons that you think are most important for your own classes. You should also consult the standards correlation table that follows when selecting chapters and lessons to include in the FlexBook for your classes.

TABLE 1.3: Standards Addressed by the Lessons in The Science of Chemistry

Lesson	California Standards	SSES Standards	AAAS Standards
1.1 The Scientific Method	1c, 1d, 1f, 1g, 1j, 1k, 1n		
1.2 Chemistry in History	1f, 1g, 1k, 1n		
1.3 Chemistry is a Science of Materials	1g, 1l, 1n		

1.1 The Scientific Method

Key Concepts

In this lesson, students explore the advancements of mankind in transportation, communication, and medicine, and gain an appreciation for scientific methods.

Lesson Objectives

- Describe the steps involved in the scientific method.
- Appreciate the value of the scientific method.
- Recognize that in some cases not all the steps in the scientific method occur, or they do not occur in any specific order.
- Explain the necessity for experimental controls.
- Recognize the components in an experiment that represent experimental controls.

Lesson Vocabulary

hypothesis A proposal intended to explain a set of observations.

theory A hypothesis that has been supported with repeated testing.

law A relationship that exists between specific observations.

experiment The act of conducting a controlled test or observations.

scientific method A method of investigation involving observation to generate and test hypotheses and theories.

superstition An irrational belief that an object, action, or circumstance not logically related to an event influences its outcome.

Strategies to Engage

- Before exploring the information in this lesson, write the following phrase on the board: “A method of thinking that allows us to discover how the world around us works.” Encourage students to focus on the key word, “discover” in the phrase. Facilitate a discussion with students about how discovery in science differs from discovery in religion, and philosophy. (In religion, discovery is based on faith/divine revelation. In philosophy,

discovery is based on logical reasoning.) Point out to students that religion, philosophy, and science attempt to discover how the world around us works. The means by which this discovery occurs varies among the three. Explain to students that in this lesson, they will learn how science makes use of scientific methods to discover how the world around us works.
- Ask students, "Have you ever walked into a room, pulled the chain to turn on a lamp, and it did not turn on?" Facilitate a discussion with students about what they would do next. (Maybe they would guess that the light bulb needs to be replaced. If replacing the light bulb does not work, maybe they would try plugging the lamp into a different outlet or plugging another appliance into the same outlet to see if there was something wrong with the outlet.) Point out to students that this scenario is an example of scientific methods at work. Explain to students how this scenario involves the use of scientific methods to generate and test hypotheses. (Developed an educated guess about the solution to the problem- hypothesis. Used controlled tests to confirm or reject the hypotheses.)

Strategies to Explore

- This lesson includes a review of the last 3,000 years in the history of human transportation, communication, and medicine. Before reading, prepare less proficient readers by having students write the following on the top of separate sheets of notebook paper:

Transportation in 1000 B.C.

Transportation in 1830

Transportation in 1995

Communication in 1000 B.C.

Communication in 1830

Communication in 1995

Medical Treatment in 1000 B.C.

Medical Treatment in 1830

Medical Treatment in 1995

As they read each section have them write key points under each heading. This will give the students a quick reference and help them to organize the information. Instruct students to write a one-paragraph summary of the information they have read in each section. DI **Less Proficient Readers**

- Have students play the game ***DAZOO***. This game is located in the Supplemental Lab Book.
- Play the ***This or That Psychic Game***, *and the* **Seven of Diamonds Game**. These games are located in the Supplemental Lab Book.

Strategies to Extend and Evaluate

- Ask students if they would describe the relationship between science and religion and/or the relationship between science and philosophy to be one of conflict, independence, dialogue, or integration. Have students support their opinions with examples from the text.
- Freeman Dyson, a noted physicist, said that the most important invention of mankind was hay. Facilitate a discussion with students about why he might have made this statement based on the readings of the first two pages.

- Read each statement in the lesson summary. Have students indicate whether or not they understand each statement by using thumb up/thumb down to show "Yes" or "No." Whenever a student uses a thumb down to show "No," use this as an opportunity to review this concept with the class. DI **English Language Learners**
- Have students read the ***Rene Blondiot and N-Rays, James Randi Versus the Dowsers***, and ***The Mysterious Bermuda Triangle*** extra readings. These readings are located in the Supplemental Workbook.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 1.1 Review Questions that are listed at the end of the lesson in the FlexBook.

Review Question with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Which of the following statements is a reasonable expression of Fleming's hypothesis?
   a. Nutrient broth kills bacteria.
   b. There are clear areas around the Penicillium mold where Staphylococcus doesn't grow.
   c. Mold kills bacteria.
   d. Penicillium mold produces a substance that kills Staphylococcus.
   e. Without mold in the culture dish, there were no clear areas in the bacteria.
2. Fleming grew Penicillium in broth, then removed the Penicillium and poured the broth into culture dishes containing bacteria to see if the broth would kill the bacteria. What step in the scientific method does this represent?
   a. Collecting and organizing data
   b. Making a hypothesis
   c. Testing a hypothesis by experiment
   d. Rejecting the old hypothesis and making a new one
   e. None of these
3. A scientific investigation is NOT valid unless every step in the scientific method is present and carried out in the exact order listed in this chapter.
   a. True
   b. False
4. Which of the following words is closest to the same meaning as ***hypothesis***?
   a. fact
   b. law
   c. formula
   d. suggestion
   e. conclusion
5. Why do scientists sometimes discard theories?
   a. the steps in the scientific method were not followed in order
   b. public opinion disagrees with the theory
   c. the theory is opposed by the church

d. contradictory observations are found
e. congress voted against it

6. Which of the following is a reasonable statement of Gary's hypothesis?

a. Different plants have different characteristics.
b. Plants that get more sunshine grow larger than plants that get less sunshine.
c. Plants that grow in the shade grow larger.
d. Plants that don't receive water will die.
e. Plants that receive the same amount of water and plant food will grow the same amount.

7. What scientific reason might Gary have for insisting that the container size for the all plants be the same?

a. Gary wanted to determine if the size of the container would affect the plant growth.
b. Gary wanted to make sure the size of the container did not affect differential plant growth in his experiment.
c. Gary want to control how much plant food his plants received.
d. Gary wanted his garden to look organized.
e. There is no possible scientific reason for having the same size containers.

8. What scientific reason might Gary have for insisting that all plants receive the same amount of water everyday?

a. Gary wanted to test the effect of shade on plant growth and therefore, he wanted to have no variables other than the amount of sunshine on the plants.
b. Gary wanted to test the effect of the amount of water on plant growth.
c. Gary's hypothesis was that water quality was affecting plant growth.
d. Gary was conserving water.
e. There is no possible scientific reason for having the same amount of water for each plant every day.

9. What was the variable being tested in Gary's experiment?

a. the amount of water
b. the amount of plant food
c. the amount of soil
d. the amount of sunshine
e. the type of soil

10. Which of the following factors may be varying in Gary's experimental setup that he did not control?

a. individual plant variation
b. soil temperature due to different colors of containers
c. water loss due to evaporation from the soil
d. the effect of insects which may attack one set of plants but not the other
e. All of the above are possible factors that Gary did not control

11. When a mosquito sucks blood from its host, it penetrates the skin with its sharp beak and injects an anti-coagulant so the blood will not clot. It then sucks some blood and removes its beak. If the mosquito carries disease-causing microorganisms, it injects these into its host along with the anti-coagulant. It was assumed for a long time that the virus of typhus was injected by the louse when sucking blood in a manner similar to the mosquito. But apparently this is not so. The infection is not in the saliva of the louse, but in the feces. The disease is thought to be spread when the louse feces come in contact with scratches or bite wounds in the host's skin. A test of this was carried out in 1922 when two workers fed infected lice on a monkey taking great care that no louse feces came into contact with the monkey. After two weeks, the monkey had NOT become ill with typhus. The workers then injected the monkey with typhus and it became ill within a few days. Why did the workers inject the monkey with typhus near the end of the experiment?

a. to prove that the lice carried the typhus virus
b. to prove the monkey was similar to man
c. to prove that the monkey was not immune to typhus

d. to prove that mosquitoes were not carriers of typhus
e. the workers were mean

12. Eijkman fed a group of chickens exclusively on rice whose seed coat had been removed (polished rice or white rice). The chickens all developed polyneuritis (a disease of chickens) and died. He fed another group of chickens unpolished rice (rice that still had its seed coat). Not a single one of them contracted polyneuritis. He then gathered the polishings from rice (the seed coats that had been removed) and fed the polishings to other chickens that were sick with polyneuritis. In a short time, the birds all recovered. Eijkman had accurately traced the cause of polyneuritis to a faulty diet. For the first time in history, a food deficiency disease had been produced and cured experimentally. Which of the following is a reasonable statement of Eijkman's hypothesis?
    a. Polyneuritis is a fatal disease for chickens.
    b. White rice carries a virus for the disease polyneuritis.
    c. Unpolished rice does not carry the polyneuritis virus.
    d. The rice seed coat contains a nutrient that provides protection for chickens against polyneuritis.
    e. None of these is a reasonable statement of Eijkman's hypothesis.

Questions 12, 13, and 14 relate to the following paragraphs.

Scientist A noticed that in a certain forest area, the only animals inhabiting the region were giraffes. He also noticed that the only food available for the animals was on fairly tall trees and as the summer progressed, the animals ate the leaves high and higher on the trees. The scientist suggested that these animals were originally like all other animals but generations of animals stretching their necks to reach higher up the trees for food, caused the species to grow very long necks.

Scientist B conducted experiments and observed that stretching muscles does NOT cause bones to grow longer nor change the DNA of animals so that longer muscles would be passed on to the next generation. Scientist B, therefore, discarded Scientist A's suggested answer as to why all the animals living in the area had long necks. Scientist B suggested instead that originally many different types of animals including giraffes had lived in the region but only the giraffes could survive when the only food was high in the trees, and so all the other species had left the area.

12. Which of the following statements is an interpretation, rather than an observation?
    A. The only animals living in the area were giraffes. B. The only available food was on tall trees. C. Animals which constantly stretch their necks will grow longer necks. D. A, B, and C are all interpretations. E. A, B, and C are all observations.

13. Scientist A's hypothesis was that
    A. the only animals living in the area were giraffes. B. the only available food was on tall trees. C. animals which constantly stretch their necks will grow longer necks. D. the animals which possess the best characteristics for living in an area, will be the predominant species. E. None of the above are reasonable statements of Scientist A's hypothesis.

14. Scientist A's hypothesis being discarded is
    A. evidence that the scientific method doesn't always work. B. a result achieved without use of the scientific method. C: an example of what happened before the scientific method was invented. D. an example of the normal functioning of the scientific method. E. an unusual case.

15. When a theory has been known for a long time, it becomes a law.
    A. True B. False

16. During Pasteur's time, anthrax was a widespread and disastrous disease for livestock. Many people whose livlihood was raising livestock lost large portions of their herds to this disease. Around 1876, a horse doctor in eastern France named Louvrier, claimed to have intvented a cure for anthrax. The influential men of the community

supported Louvrier's claim to have cured hundreds of cows of anthrax. Pasteur went to Louvrier's hometown to evaluate the cure. The cure was explained to Pasteur as a multi-step process during which: 1) the cow was rubbed vigorously to make her as hot as possible; 2) long gashes were cut into the cows skin and turpentine was poured into the cuts; 3) an inch-thick coating of cow manure mixed with hot vinegar was plastered onto the cow and the cow was completely wrapped in a cloth. Since some cows recover from anthrax with no treatment, performing the cure on a single cow would not be conclusive, so Pasteur proposed an experiment to test Louvrier's cure. Four healthy cows were to be injected with anthrax microbes, and after the cows became ill, Louvrier would pick two of the cows (A and B) and perform his cure on them while the other two cows (C and D) would be left untreated. The experiment was performed and after a few days, one of the untreated cows died and one of them got better. Of the cows treated by Louvrier's cure, one cow died and one got better. In this experiment, what was the purpose of infecting cows C and D?

A. So that Louvrier would have more than two cows to choose from. B. To make sure the injection actually contained anthrax. C. To serve as experimental controls (a comparison of treated to untreated cows). D. To kill as many cows as possible.

17. A hypothesis is

A. a description of a consistent pattern in obervations. B. an observation that remains constant. C. a theory that has been proven. D. a tentative explanation for a phenomenon.

18. A scientific law is

A. a description of a consistent pattern in obervations. B. an observation that remains constant. C. a theory that has been proven. D. a tentative explanation for a phenomenon.

19. A number of people became ill after eating oysters in a restaurant. Which of the following statements is a hypothesis about this occcurence?

A. Everyone who ate oysters got sick. B. People got sick whether the oysters they ate were raw or cooked. C. Symptoms included nausea and dizziness. D. The cook felt really bad about it. E. Bacteria in the oysters may have caused the illness.

20. Which statement best describes the reason for using experimental controls?

A. Experimental controls eliminate the need for large sample sizes. B. Experimental controls eliminate the need for statistical tests. C. Experimental controls reduce the number of measurements needed. D. Experimental controls allow comparison between groups that are different in only one independent variable.

21. A student decides to set up an experiment to determine the relationship between the growth rate of plants and the presence of detergent in the soil. He sets up 10 seed pots. In five of the seed pots, he mixes a precise amount of detergent with the soil and the other five seed pots have no detergent in the soil. The 5 seed pots with detergent are placed in the sun and the five seed pots with no detergent are placed in the shade. All 10 seed pots receive the same amount of water and the same number and type of seeds. He grows the plants for two months and charts the growth every two days. What is wrong with his experiment?

A. The student has too few pots. B. The student has two independent variables. C. The student has two dependent variables. D. The student has no experimental control on the soil.

22. A scientist plants two rows of corn for experimentation. She puts fertilizer on row 1 but does not put fertilizer on row 2. Both rows receive the same amount of sun and water. She checks the growth of the corn over the course of five months. What is acting as the control in this experiment?

A. Corn without fertilizer. B. Corn with fertilizer. C. Amount of water. D. Height of corn plants.

23. If you have a control group for your experiment, which of the following is true?

A. There can be more than one difference between the control group and the test group, but not more three differences or else the experiment is invalid. B. The control group and the test group may have many differences between them. C. The control group must be identical to the test group except for one variable. D. None of these are true.

24. If the hypothesis is rejected by the experiment, then:
    A. the experiment may have been a success. B. the experiment was a failure. C. the experiment was poorly designed. D. the experiment didn't follow the scientific method.

25. A well-substantiated explanation of an aspect of the natural world is a:
    A. theory. B. law. C. hypothesis. D. None of these.

1.2 Chemistry in History

Key Concepts

In this lesson, students explore the definition and history of chemistry, the Law of Conservation of Mass, and the use of scientific methods.

Lesson Objectives

- Give a brief history of how chemistry began.
- State the Law of Conservation of Mass.
- Explain the concept of a model, and create simple models from observations.

Lesson Vocabulary

hypothesis A proposal intended to explain a set of observations.

theory A hypothesis that has been supported with repeated testing.

law A relationship that exists between specific observations.

scientific method A method of investigation involving observation to generate and test hypotheses and theories.

chemistry The science of the composition, structure, properties, and reactions of matter.

Strategies to Engage

- Before beginning the lesson, ask students to predict which of the following two statements are true and which statement is false:

a. Chemistry began as the quest for a way to transform common metals into gold. (True)

b. "Chemistry" was derived from an Arabic word. (True)

c. New matter is formed in chemical reactions. (False)

Ask students to make their predictions based on what they already know. Have a volunteer who answered correctly that the first two statements are true and the last statement false explain how they came up with their answer.

Strategies to Explore

- As you explore the section entitled, "The Origins of Chemistry Was Multicultural," have students write down what they believe to be the main idea of each paragraph. Instruct each student to pair up with another student and come to a consensus as to what they believe to be the main idea. Have each pair of students team with another pair, so that they are in groups of four and again, come to a consensus. Have each group of students share results with the class. DI **Less Proficient Readers**
- As you explore the section entitled, "The Origins of Chemistry Was Multicultural," students will come across the term "quantitative." Students often have trouble telling the difference between quantitative (numerical) data and qualitative (descriptive) data. Have students observe Figure 1.16. Instruct students to come up with three examples of qualitative data about the man in the picture. He is tall, wearing blue pants, and smells good, etc. Instruct student to come up with three examples of quantitative data about the man in the picture. He weighs 180 lbs., is 5'7" tall, and his body temperature is $98.6°C$.
- Demonstrate the law of conservation of mass by pouring 15 mL of vinegar into an empty bottle. Pour about 5 grams of baking soda into a balloon. Place the balloon onto the top of the bottle being careful not to allow any of the baking soda to fall inside of the bottle. Obtain the mass of the soda bottle and balloon. Allow the baking soda to fall into the vinegar. After the reaction has occurred, obtain the mass of the soda bottle and balloon. Explain to the students that, according to the law of conservation of mass, in an ordinary chemical reaction, matter is not created nor destroyed, but may change form.

Strategies to Extend and Evaluate

- Robert Boyle is often called the father of modern chemistry. This honor is also sometimes given to Antoine Lavoisier. Choose a few students and have them debate which chemist should be regarded as the father of modern chemistry. Students should be prepared to defend their choices and try to convince the remaining students that the chemist is the father of modern chemistry. At the end of the debate, have the students vote on which group defended their chemist better.
- Read each statement in the lesson summary. Have students indicate whether or not they understand each statement by using thumb up/thumb down to show "Yes" or "No." Whenever a student uses a thumb down to show "No," use this as an opportunity to review this concept with the class. DI **English Language Learners**

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 1.2 Review Questions that are listed at the end of the lesson in the FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Where does the word "chemistry" come from?
2. Consider the following data about John's study habits, and grades:

a. Propose a qualitative (words, but no math) model that might describe how the length of time John spends studying relates to how well he does on the test?
b. If John wants to earn 92% on his next test, should he study for about 6 hours, 9 hours, 12 hours, or 18 hours? Justify your answer.
c. If John studies for 7 hours, do you think he will score 15%, 97%, 68%, or 48%? Justify your answer.

3. Helen wanted to know if lemon juice chemically reacts with tea to lighten its color. So Helen added 25 drops of lemon juice to 250 mL of tea and observed that the tea colored lightened significantly. Helen wanted to make sure that the color lightening was the result of a chemical reaction and not the result of dilution. Which one of the following activities should Helen carry out to serve as a control for this experiment?
    a. Helen should add 25 drops of orange juice to another 250 mL sample of tea.
    b. Helen should add 25 drops of distilled water to another 250 mL sample of tea.
    c. Helen should add 25 drops of lemon juice to a 250 mL sample of distilled water.
    d. Helen should add 25 drops of tea to a 250 mL sample of lemon juice.
    e. Helen should add 25 drops of tea to a 250 mL sample of tea.

1.3 Chemistry is a Science of Materials

Key Concepts

In this lesson, students explore the role of a chemist as a scientist who studies the properties of matter.

Lesson Objectives

- Give examples of chemical properties a scientist might measure or observe in a laboratory.
- Explain the difference between a physical change and a chemical change, giving examples of each.
- Identify the situations in which mass can be converted to energy and energy can be converted to mass.

Lesson Vocabulary

alloy A solution (or a special kind of mixture), in which at least one of the components is a metal.

physical change Changes that do not alter the identity of the substance.

chemical change A change that occurs when one substance is turned into an entirely new substance as a result of a chemical reaction.

Strategies to Engage

- Have students observe Figure 3. Facilitate a discussion with students about how everyday life would be different without plastics. Explain to students that plastics are just one of many products that came about through scientists' attempts to control the properties of matter in order to use them to our advantage.

Strategies to Explore

- Have students read the lesson objectives. Instruct students to create a five-question quiz from those three objectives. Have each student exchange quizzes with a classmate. As students explore this lesson, instruct them to answer the five questions. At the end of the lesson, have them give the quiz back to the original student who will grade the quiz. Encourage students to discuss discrepancies.
- Demonstrate the difference between chemical and physical changes using a birthday candle. Use a Bunsen burner, lighter, or match to melt one end of a candle. Allow students to observe the melting candle and the melted candle wax that results. Then, light the candlewick and allow it to burn. Facilitate a discussion with

students about the difference between melting and burning the candle. Students should notice that, in the case of melting, the wax was the same substance as the candle. If students mention that when the candle was burned it "disappeared", inform them that it did not "disappear". Rather it was changed into carbon dioxide gas and water vapor. Emphasize to students that when a substance undergoes a physical change, as was the case when the candle melted, no new substances are produced. On the other hand, when a substance undergoes a chemical change, as was the case when the candle burned, new substances are formed.
- Have students complete the lab ***Candle Observation***. This lab is located in the Supplemental Lab Book.

Strategies to Extend and Evaluate

- Have interested students participate in a mock trial in which plastics are the defendants. Have a team of student-lawyers defend plastics and another group of student-lawyers prosecute plastics. The remainder of the class will serve as the jury. Encourage students to focus their arguments on the benefits and consequences of plastics on society and the environment.
- Challenge interested students to choose a material such as paper, sugar, or water. Instruct them to write up methods to demonstrate the material undergoing either a physical change or a chemical change, and if possible, perform their demonstration for the class. Have the class determine whether each change is physical or chemical.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 1.3 Review Questions that are listed at the end of the lesson in the FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Name the two types of changes that chemists are primarily interested in.
2. Decide whether each of the following statements is true or false.
   a. Plastics were developed in Rome around 300 AD
   b. Bronze is an example of an alloy
   c. Plastic is an example of an alloy
   d. Brass is an example of an alloy
3. Match the following alloys with their common names.
   a. Bronze - a. tin and copper
   b. Brass - b. is not an alloy
   c. Plastic - c. iron and carbon
   d. Tin - d. copper and zinc
   e. Steel - e. is not an alloy
4. Decide whether each of the following statements is true or false.
   a. Physical changes are typically accompanied by a color change

b. A burning campfire is an example of a chemical change
c. When you heat your house with coal, the coal undergoes a chemical change
d. When you drop a plate, and it breaks, the plate undergoes a physical change

5. In each of the following examples, determine whether the change involved is a physical change or a chemical change.
    a. Flattening a ball of silly putty
    b. Combining a bowl of cherries and a bowl of blueberries
    c. Boiling water
    d. Cooking an egg
6. Judy has two beakers filled with clear liquids, and she needs to know whether the liquid in the first beaker is the same as the liquid in the second beaker. In which scenario does Judy use physical properties to answer her question, and in which scenario does Judy use changes in chemical properties to answer her question?
    a. Judy smells the two liquids and notices that the liquid in the first beaker has a strong odor, while she can't smell the liquid in the second beaker at all.
    b. Judy mixes some table salt into the first beaker and notices that a white precipitate forms. She then mixes some table salt into the second beaker, but nothing happens.

1.4 Matter

Key Concepts

In this lesson, students explore the definition and composition of matter. Students will also explore the difference between mass and weight.

Lesson Objectives

- Define matter and explain how it is composed of building blocks known as "atoms."
- Distinguish between mass and weight.

Lesson Vocabulary

matter Anything of substance that has mass and occupies space.

atom The basic building block of all matter. There are 117 known types of atoms. While atoms can be broken down into particles known as electrons, protons and neutrons, this is very difficult to do.

element A type of atom. There are 117 known elements.

molecule Two or more atoms bonded together. Specific molecules, like water, have distinct characteristics.

Periodic Table A way of summarizing all the different atoms that scientists have discovered. Each square in the periodic table contains the symbol for one of the elements.

mass An intrinsic property of matter that can be used to measure the quantity of matter present in a sample.

weight A measurement of how strongly gravity pulls on an object.

Strategies to Engage

- Introduce lesson concepts by asking students to observe Figure 1.25 and recall what they know about matter. Guide them in focusing their prior knowledge.

Ask: What are some things that all objects have in common? (All objects are composed of matter.)

Ask: How do you know that an ant is composed of matter? (It has mass and takes up space.)

Ask: Name some examples of "things" that are not composed of matter. (Emotions, senses, ideas.)

Ask: How do you know that these things are not composed of matter? (They do not have mass and do not take up space.)

Strategies to Explore

- Facilitate a discussion with students about the relationship between building materials and atoms. Ask: If building materials are like atoms, what are elements? (The elements would be the types of building materials such as the bricks, wood, and the insulation.)
- Write the following chemical formulas on the board: CO, CO_2, C_2H_4, $CaCO_3$, and CN. Ask: What do these chemical formulas have in common? (They all contain the element carbon. Point out to students that the one element, carbon, is present in all five of these chemical formulas. Explain to students that all compounds are made from elements and that elements such as carbon can combine with other elements to form compounds.)
- Explain to students that the relationship between mass and weight is given by the equation $W = mg$. Where "W" represents weight in Newtons, "m" represents mass in kilograms, and "g" represents acceleration due to gravity. Have students find out their weight on other planets at http://www.exploratorium.edu/ronh/weight/

Strategies to Extend and Evaluate

- Have each student record the four sentences in this section that most clearly represent the main ideas. Read key sentences in the text and have students raise their hands if they have recorded that sentence. Facilitate a discussion in which students defend their selections. DI **Less Proficient Readers**
- Ask students to search for examples of the terms "mass" and "weight" being used incorrectly. Have them quote the claim, reference the source, and then explain what is wrong.

Lesson Worksheets

Copy and distribute the worksheet in the *CK-12 Chemistry Workbook* named ***Mass versus Weight***. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 1.4 Review Questions that are listed at the end of the lesson in the FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. What is matter?
2. In this chapter, we'll learn about atoms, which are the building blocks of all matter in the universe. As of 2007, scientists only know of 117 different types of atoms. How do you think it's possible to generate so many different forms of matter using only 117 types of building blocks?
3. Which do you think has more matter, a cup of water or a cup of mercury? Explain.
4. Decide whether each of the following statements is true or false.

a. Mass and weight are two words for the same concept.
b. Molecules are bonded together to form atoms.
c. Alchemists couldn't make gold out of common metals because gold is an element.
d. The symbol for Gold in the periodic table is Gd.

5. Would you have more mass on the moon or on Earth?
6. Would you have more weight on the moon or on Earth? The force of gravity is stronger on the Earth than it is on the moon.
7. Match the following terms with their meaning.

a. Mass - a. a measure of the total quantity of matter in an object
b. Volume - b. a measure of how strongly gravity pulls on an object
c. Weight - c. a measure of the space occupied by an object

8. For the following statements, circle all of the options that apply: Mass depends on...

a. the total quantity of matter
b. the temperature
c. the location
d. the force of gravity

Volume depends on...

a. the total quantity of matter
b. the temperature
c. the object's shape (independent of size)
d. the object's size (independent of shape)

Weight depends on...

a. the total quantity of matter
b. the temperature
c. the location
d. the force of gravity

1.5 Energy

Key Concepts

In this lesson, students explore the definition and some forms of energy. Students will also explore the Law of Conservation of Matter and Energy.

Lesson Objectives

- Define heat and work.
- Distinguish between kinetic energy and potential energy.
- State the Law of Conservation of Matter and Energy.

Lesson Vocabulary

heat Energy that is transferred from one object to another object due to a difference in temperature. Heat naturally flows *from* a hot object *to* a cooler object.

force Any push or pull.

work A force applied over a distance.

kinetic energy Energy associated with motion.

potential energy Stored energy. Potential energy depends on an object's position (or mixture's composition).

chemical potential energy Potential energy stored in the atoms, molecules, and bonds of matter.

Law of Conservation of Energy Energy cannot be created or destroyed; it can only be changed from one form to another.

Law of Conservation of Mass and Energy The total amount of mass and energy in the universe is conserved.

Strategies to Engage

- Have students read the lesson objectives. Ask students to write down and try to complete each objective. Instruct students to use a scale of 1-5 (1= not sure, 5 = very sure) to record how sure they are that they have correctly completed each objective. As you explore this lesson, encourage students to change their answers as necessary.

Strategies to Explore

- Place 100 *mL* of glycerin into a beaker. Have one student-volunteer obtain the temperature of the glycerin in the beaker. Have another student-volunteer use a metal spoon to stir the glycerin in the beaker for about 40 *seconds*. Have a third student-volunteer use a thermometer to obtain the temperature of the glycerin after it has been stirred. Explain to students that this demonstration shows that energy can be transferred as heat or work. Work is a force applied over a distance. When the student stirred the glycerin, work was done on the glycerin. Thermal energy was transferred from the particles of the glycerin to the thermometer in the form of heat. Heat is simply energy that is transferred from an object with a higher temperature to an object with a lower temperature.
- Have students complete the lab named ***Energy Lab- Recognizing PE***. This lab is located in the Supplemental Lab Book.

Strategies to Extend and Evaluate

- Have students write a one-paragraph summary of this lesson. Instruct students to correctly use the following terms in their paragraph: energy, kinetic, potential, transfer, heat, work, and temperature.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 1.5 Review Questions that are listed at the end of the lesson in the FlexBook.

Review Questions with Sample Answers

1. Classify each of the following as energy primarily transferred as heat or energy primarily transferred as work:
   a. The energy transferred from your body to a shopping cart as you push the shopping cart down the aisle.
   b. The energy transferred from a wave to your board when you go surfing.
   c. The energy transferred from the flames to your hotdog when you cook your hotdog over a campfire.
2. Decide whether each of the following statements is true or false:
   a. When heat is transferred to an object, the object cools down.
   b. Any time you raise the temperature of an object, you have done work.
   c. Any time you move an object by applying force, you have done work.
   d. Any time you apply force to an object, you have done work.
3. Rank the following scenarios in order of *increasing* work:
   a. You apply 100 N of force to a boulder and successfully move it by 2 m.
   b. You apply 100 N of force to a boulder and successfully move it by 1 m.
   c. You apply 200 N of force to a boulder and successfully move it by 2 m.
   d. You apply 200 N of force to a boulder but cannot move the boulder.
4. In science, a vacuum is defined as space that contains absolutely no matter (no molecules, no atoms, etc.) Can energy be transferred as heat through a vacuum? Why or why not?

5. Classify each of the following energies as kinetic energy or potential energy:
   a. The energy in a chocolate bar.
   b. The energy of rushing water used to turn a turbine or a water wheel.
   c. The energy of a skater gliding on the ice.
   d. The energy in a stretched rubber band.
6. Decide which of the following objects has more kinetic energy:
   a. A 200 lb. man running at 6 mph or a 200 lb. man running at 3 mph.
   b. A 200 lb. man running at 7 mph or a 150 lb. man running at 7 mph.
   c. A 400 lb. man running at 5 mph or a 150 lb. man running at 3 mph.
7. A car and a truck are traveling along the highway at the same speed.
   a. If the car weighs 1500 kg and the truck weighs 2500 kg, which has more kinetic energy, the car or the truck?
   b. Both the car and the truck convert the potential energy stored in gasoline into the kinetic energy of motion. Which do you think uses more gas to travel the same distance, the car or the truck?
8. You mix two chemicals in a beaker and notice that as the chemicals react, the beaker becomes noticeably colder. Which chemicals have more chemical potential energy, those present at the start of the reaction or those present at the end of the reaction?

Chapter 1 Enrichment

Extra Readings

Rene Blondlot and N-Rays

In 1903, Rene Blondlot was a distinguished professor of physics at the University of Nancy, France. He was a member of the French Academy of Sciences and had won several scientific awards. He had designed and carried out a brilliant experiment to measure the speed of electricity traveling through a conductor. Other scientists duplicating Blondlot's methods found that the method worked and they got the same result as Blondlot. Thus, his results were verified.

Later in his career, while trying to polarize x-rays, Blondlot claimed to have discovered a new invisible radiation similar to x-rays which he called N-rays (after the town of Nancy). Blondlot claimed that N-rays were emitted by all substances except wood and he detected them with an instrument he designed using an aluminum prism to scatter the rays and a fluorescent thread to detect the rays. Fourteen of Blondlot's friends (also scientists) confirmed the existence of the N-rays. Some other scientists tried Blondlot's experimental set up and agreed with his results. The French Academy of Science was preparing to award Blondlot their highest prize, the LaLande Prize.

Dr. Robert Wood, an American scientist, attempted to reproduce Blondlot's experiment in his own lab. Not only was Wood unable to obtain Blondlot's results but some of the observations reported by Blondlot seemed to Wood to be impossible. *Nature* magazine was skeptical of Blondlot's result because other scientists in England and Germany were also unable to duplicate Blondlot's result. The magazine sent Dr. Wood to investigate Blondlot's discovery.

Wood visited Blondlot's lab and asked for a demonstration of the experiment. Wood looked through the eyepiece of the instrument but saw no effect of N-rays on the thread. He was told by Blondlot that his eyes weren't properly sensitized. At one point, while Blondlot and his assistant were operating the instrument, Wood secretly reached into the machine and removed the prism. Both Blondlot and his assistant, however, continued to "see" the evidence of N-rays when it was impossible for the instrument to work. Wood then tried to secretly put the prism back in place,

but the assistant saw him and thought that Wood was removing the prism. The next time they ran the experiment, neither Blondlot nor the assistant could see any N-rays, even though the machine was in proper working order.

Wood published the results of his visit to Blondlot's lab and the contentions of Blondlot, his assistant, and colleagues was discredited. The French Academy of Science had already published over 100 papers about N-rays. The Academy went ahead and awarded the LaLande prize to Blondlot but it was presented as rewarding his entire career and no mention was made of N-rays. Ten years later, all mention of N-rays had been removed from French science books and French encyclopedias.

There are other similar stories - some about scientists who made serious errors in experiments and others about scientists who faked data - if you are interested in reading about some, you could search internet stories on *polywater* or *cold fusion*.

James Randi versus the Dowsers

Dowsing is the process of using wooden twigs or metal rods to locate hidden water or pieces of metal. The dowser holds the sticks or rods in his hands and when they swing together, it indicates the presence of water or metal.

Starting Position of Rods — Position of Rods when Locating

With the rods in the starting position, the dowser walks across a search area and when he/she passes over underground water or a hidden piece of metal, the rods will swing together indicating the presence of water or metal.

James Randi, a former magician, who now spends his time debunking paranormal charlatans, has made a standing offer, originally $10,000 but now $1,000,000, to anyone who can pass controlled tests to prove they have paranormal or supernatural powers (this includes dowsing). Mr. Randi uses what are called double blind experiments to test the claims of dowsers. A double blind test requires that neither the dowser nor the judges know the position of the dowsers search object. Over 1,000 people have attempted to claim the prize money, none have succeeded.

Consider the case of Stanley Wojcik, who claimed to be an expert dowser who could locate hidden pieces of metal in over 90% of his trials. Mr. Wojcik supplied reference letters from individuals who supported his claims. Mr. Wojcik's dowsing rods were two coat hangers straightened out to form L-shaped pieces. His procedure was to proceed forward with the rods projecting straight out in front of him until some object was "sensed" and then the rods would swing together.

The test began with Mr. Wojcik being asked to locate a small pile of coins placed on a table in plain sight. This was done to show the judges how the dowsing rods behaved when locating the hidden object and to offer proof that there was nothing in the location to inhibit the function of the dowsing rods. It is common for dowsers who fail the test to offer excuses for the failure . . . the most common excuse is that there is something in the location that interfered with the test . . . like water pipes in the floor or something of that sort.

Mr. Wojcik walked around the room with the dowsing rods extended and when he reached the coins on the table, the dowsing rods came together. Then Mr. Wojcik was asked if the rods would still work if the coins were placed in an envelope and Wojcik replied in the affirmative. During the second test, when the coins were placed in an envelope and placed on the same table in the same place as before, the rods again came together precisely over the envelope. In the next dozen tests, nine more envelopes identical to the first but containing small lumps of paper to match the lump caused by the coins were placed around the room. Even though the odds would indicate that the dowser would correctly find the envelope containing the coins once in ten tries, Mr. Wojcik failed to find the coins even one time. When Mr. Wojcik indicated the test was flawed because of water pipes, the test was moved to another room and Mr. Wojcik still failed every time. The dowser had scored 100% on the trials where he could see the object and 0% on the blind trials.

Tests such as this have been performed with many dowsers over the years, but Mr. Randi still has the $1,000,000.

Checking the Data on the "Mysterious" Bermuda Triangle

The "Bermuda Triangle" is a triangular expanse of ocean between the three vertices of Bermuda, Puerto Rico, and Miami. The "mystery" of the Bermuda Triangle was set forth in a series of three books written by Charles Berlitz in the 1970's. Since the publication of the three books, dozens of other books, articles, stories, and several TV movies about the Bermuda Triangle have appeared.

Berlitz's books contained a collection of stories of boats, airplanes, and people mysteriously lost at sea in the Bermuda Triangle, and claimed that all the stories were true and that they offered proof that there was something strange about the Triangle. Berlitz convinced millions of people that there was some unknown force in the Triangle that caused planes, boats, and people to disappear. This unknown force has variously been attributed to a sunken flying saucer, the lost city of Atlantis, or some distortion in the earth's magnetic field.

When an unbelievable story is claimed to be true, the best place to start checking the story is to examine the data upon which the hypothesis is based.

Berlitz identified approximately 80 incidents that he claimed occurred in the Bermuda Triangle. Subsequent authors have stated that there are thousands of such incidences but do not identify any of them. Skeptics who investigated the original 80 incidents have determined that 41 of them did not occur at all. That is, there was never an airplane or boat by the name given in the story; there was never a report made to the U.S. Navy, Coast Guard; or to any police

department; the people named cannot be located by the names given in the story; and there were no flight plans or travel plans filed at the airport or harbor of origination. It is presumed, therefore, that these are fictional incidents.

Of the total number of incidents claimed, only 39 have any evidence indicating that they actually took place. Of the 39, 10 were accidents in which a ship was found abandoned. Bermuda Triangle authors indicate that the people disappeared with no explanation. For these 10 cases, however, the crews were rescued and produced quite normal explanations of what happened and whly the ship was abandoned.

The other 29 incidents are indicated on the map.

Of the original 80 incidents, 41 were fictitious, 10 turned out to be quite normal, and of the 29 remaining incidents, only 4 of them actually occurred inside the Triangle. As you can see on the map, one incident occurred in the Gulf of California, over $2,000$ *miles* from the Bermuda Triangle. You can also see three incidents that occurred on the European side of the Atlantic Ocean, also over $2,000$ *miles* from the Triangle. These three occurred off the coast of Ireland, off the coast of Portugal, and near the Azores Islands. It is absurd to include these events in any examination of the Bermuda Triangle.

Berlitz reported incidents in his book using language and shortage of details to make the incidents seem as mysterious as possible. More complete reports often remove the mystery.

Example report by Berlitz: "Thirty-nine persons vanished north of the Triangle on a flight to Jamaica on February 2, 1953. An SOS, which ended abruptly without explanation, was sent by the ***British York*** transport just before it disappeared. No trace was ever found."

A more complete report: The flight plan of the transport was to fly from the Azores (near Portugal) to Newfoundland, Canada. After an overnight stopover, the plane was to continue on to Jamaica the following day. On the flight to Canada, the plane encountered strong winds up to 75 *miles* per hour and torrential rains in the mid-Atlantic. The

crew sent an SOS which ended abruptly and no parts of the airplane were found.

This airplane did not crash in the Bermuda Triangle, in fact, it never even flew through the Bermuda Triangle. The only connection to the Triangle was a future flight plan. When metal airplanes fall into the ocean, they sink . . . and when they sink, radio messages cease abruptly. Berlitz uses words like "vanished" and "disappeared" rather than "crashed in the ocean and sank" to make the incident seem mysterious.

Another Berlitz report: Berlitz's description of the loss of Eastern Flight 401 indicated that while flying through the Triangle, the Eastern flight "suffered a loss by disintegration". This description would lead us to believe that the flight was somewhere in the Triangle when suddenly pieces of the airplane began to fall off for no apparent reason.

Surviving crew member's report: The crew members reported that while over the Florida Everglades (not in the Triangle), they turned off the autopilot and while trying to fix a navigation problem, failed to notice a loss in altitude. The plane flew into the ground and "disintegrated". End of mystery.

Another Berlitz report: This incident involves Christopher Columbus. Berlitz reports in his book (and quotes from Columbus' logbook) that Columbus wrote about a "fireball which circled his flagship".

Other readers of the logbook report: Columbus wrote of "a great flame which fell into the sea." There is no indication or implication in Columbus' logbook that the flame circled his ship. A meteor burning through the atmosphere is a spectacular sight and somewhat rare. One that flew around in a circle would indeed be a mystery.

The primary incident that Bermuda Triangle enthusiasts would point to is an incident that occurred on December 15, 1945. According to Berlitz, five fully equipped Avenger torpedo bombers took off from Fort Lauderdale Naval Air Station on a flight into the Triangle and back. At the time the planes should have returned, the flight leader reported over the radio they were lost and confused about directions. About 45 $minutes$ later, the planes vanished from radar screens. A rescue plane sent to find them also disappeared. No trace was found of either flight. Berlitz also reported some strange radio transmissions by the pilots and flight leader.

The official Naval report of the incident, however, again, is a somewhat different story. The flight was a training flight for new pilots. Only the flight leader's compass was turned on. During the flight, the flight leader's compass failed and this failure was reported by radio. The failure was discovered after the planes were dangerously low on fuel. The pilots were lost and flew around in confusion until out of fuel and then fell into the sea. The rescue plane that was sent was known to be a dangerous plane because it leaked gas fumes inside the plane. A ship in the area observed this rescue plane explode and fall into the sea. Records and witness accounts of the radio transmissions from the flight show no strange or unusual transmissions. From the official report, there appears to be no mystery. The media, however, preferred the mysterious scenario and chose not to check the facts.

The desire to see favorable results where none exist is the source of much of the "data" presented by supporters of the paranormal. There are also examples of this failing in orthodox science. To protect ourselves from such wrong-headed thinking, we must always be skeptical and when we suspect flawed procedures, CHECK THE DATA AND HOW IT WAS COLLECTED.

CHAPTER 2 Chemistry – A Physical Science TE

Chapter Outline

Chemistry – A Physical Science

Outline

The chapter *Chemistry – A Physical Science* consists of four lessons that cover measurement and the mathematics of measurement and formulas.

- **Lesson 2.1 Measurements in Chemistry**
- **Lesson 2.2 Using Measurements**
- **Lesson 2.3 Using Mathematics in Chemistry**
- **Lesson 2.4 Using Algebra in Chemistry**

Overview

In these lessons, students will explore:

- The units used to express mass, volume, length, and temperature.
- Metric prefixes, scientific notation, and significant figures.
- The use of dimensional analysis and significant figures in chemistry problem solving.
- The use of algebra in chemistry problem solving.

Science Background Information

- The Metric System

In the late 18th century, Louis XVI of France charged a group of scientists to reform the French system of weights and measures. It was widely recognized at the time that it was an inconsistent and disorganized collection of measurements that varied with location and often on obscure bases. Providing a scientifically observable system with decimally based divisions was the charge assigned to a group from the French Academy of Sciences, which included

Pierre Simon LaPlace and J.J. Lagrange. They sought to create bases of measurement linked to the scientifically verifiable values such as the Earth's circumference.

The unit of length, defined as a meter, was introduced in 1791 after careful measurement of the Earth's radius and the recognition that the planet was not perfectly spherical but instead possessed an oblate spheroid shape. The meter was designated as one ten-millionth of the length of the Earth's meridian through the city of Paris from the North Pole to the Equator.

For the measurement of volume, the SI unit devised in 1795 was the cubic meter, which was based on the volume of a cube with sides of one meter each. The large size of this unit has largely resulted in the more common use of the smaller metric unit of the liter, defined as 0.001 cubic meters.

The kilogram was settled upon in 1799 as the mass standard, based on the value of a platinum bar. Now the contemporary standard for the kilogram is stored at the Bureau International des Poids et Mesures (BIPM) in Sevres, France as a Platinum-iridium alloy.

The original definition of the principal time unit, the second was considered to be $\frac{1}{86,400}$th of the mean solar day. Due to inconsistencies in the rate of the Earth's rotation, the modern definition is linked to the radiation correlating to the orbital transitions of the cesium -133 isotope.

Since the 1960s, the International System of Units has been internationally agreed upon as the standard metric system.

- What is the Kelvin Temperature Scale?

There are three different temperature scales in use in the world today. Mainly the United States utilizes the Fahrenheit scale, which was introduced by Daniel Gabriel Fahrenheit in 1724. The non-intuitive reference points on the Fahrenheit system ($212°F$ and $32°F$) for the boiling and freezing points of water, respectively) are replaced in the more universally accepted Celsius, or Centigrade system, devised by Anders Celsius in 1742, by $100°C$ and $0°C$ for scientific applications, however, both scales are inconveniently constructed in that a substantial portion of the scale consists of negative values for temperature. For many physical considerations, the use of a Celsius or Fahrenheit temperature that is a negative number produces an impossible result, such as in the Ideal Gas Law, $(pV = nRT)$.

In 1848, William Thomson Kelvin, a British physicist proposed the scale that is now named in his honor. In the design of this system, there are no negative values for temperature with the lowest value on the scale known as absolute zero. Substances at this theoretical point would display a complete absence of kinetic energy, thus atoms at absolute zero would cease all motion.

The Kelvin and Celsius scales are routinely used in chemical measurements and are conveniently constructed in that temperature change between any two points are exactly the same. Most laboratory thermometers available today are graduated in the Celsius system yet transition to the accepted SI Kelvin units is straightforward; since $0K = -273.15°C$, adding 273.15 *degrees* to the Celsius temperature will yield the correct Kelvin value. Note that because the Kelvin system is an absolute scale, the degree symbol (°) is omitted in reporting the Kelvin temperature.

Pacing the Lessons

Use the table below as a guide for the time required to teach the lessons of *Chemistry-A Physical Science.*

TABLE 2.1: Class Periods per Lesson

Lesson	Number of 60 Minute Class Periods
2.1 *Measurements in Chemistry*	1.0
2.2 *Using Measurements*	2.0
2.3 *Using Mathematics in Chemistry*	1.0

TABLE 2.1: (continued)

Lesson	Number of 60 Minute Class Periods
2.4 *Using Algebra in Chemistry*	1.0

Managing Materials

The following materials are needed to teach the strategies and activities described in the Teachers Edition of the FlexBook for *Chemistry-A Physical Science.*

TABLE 2.2: **Chemistry-A Physical Science Materials List**

Lesson	Strategy or Activity	Materials Needed
Lesson 2.1	Exploration Activity	index cards
Lesson 2.2	Metric Scavanger Hunt	rulers, balances, meter sticks, graduated cylinders
Lesson 2.3		
Lesson 2.4		

Multimedia Resources

You may find these additional Web-based resources helpful when teaching *Chemistry-A Physical Science.*

- Introduction to the Metric System http://videos.howstuffworks.com/hsw/5890-scientific-method-the-metric-system.htm
- Metric Conversion http://www.tecc.tv/oldmhh/manthhomeworkhotline.com/old/metric.html
- Metric Equivalents http://www.harcoutschool.com/activity/con_math/g03c25.html
- Comparing and Ordering Numbers in Scientific Notation http://www.learnalberta.ca/content/mejhm/html/video_interactive/exponents/exponentsInteractive.html

Possible Misconceptions

Making the FlexBook Flexible

An important advantage of the FlexBook is the ability it gives you, the teacher, to select the chapters and lessons that you think are most important for your own classes. You should also consult the standards correlation table that follows when selecting chapters and lessons to include in the FlexBook for your classes.

TABLE 2.3: **Standards Addressed by the Lessons in Chemistry-A Physical Science**

Lesson	California Standards	SSES Standards	AAAS Standards
Lesson 2.1	1a, 4e, 4f, 4g		
Lesson 2.2	1a		
Lesson 2.3	1a		
Lesson 2.4	1a, 1e		

2.1 Measurements in Chemistry

Key Concepts

In this lesson, students explore the units used to express mass, volume, length, and temperature.

Lesson Objectives

- State the measurement systems used in chemistry.
- State the different prefixes used in the metric system.
- Do unit conversions.
- Use scientific notation and significant figures.
- Use basic calculations and dimensional analysis.
- Use mathematical equations in chemistry.

Lesson Vocabulary

International System of Units, SI The SI system of units is the modern form of the metric system and is generally a system devised around the convenience of multiples of 10.

Kelvin temperature scale The kelvin is unit of increment of temperature and is one of the seven SI basic units. The Kelvin scale is thermodynamic absolute temperature scale where absolute zero is the absence of all thermal zero. At $K = 0$, there is no molecular motion. The kelvin is not referred to as a "degree", it is written simply as K, not $^\circ K$.

Strategies to Engage

- Point to an item in the room and say to students "Do you think that (item) is 10?". If students reply "10 what?", ask them to list some measurements to which the "10 could possibly refer. Inches, meters, kilograms, age, etc. Explain to students that measurements without numbers are meaningless. Inform students that in this lesson, they will explore various measurement units.
- Inform students that On September 23, 1999 NASA lost its $125 *million* Mars Climate Orbiter. Review findings indicate that one team used English units of measurement while another team used metric units. Facilitate a discussion with students about the importance of having and using a standardized measurement system.

Strategies to Explore

- Hand each group of three students an index card. Inform students that the first group to construct a box (without a lid) that will hold exactly 1.00 *mL* of water will win a prize. (A box that is 1 *cm* on each edge will have a volume of 1 cm^3, which equals 1 *mL*.)
- Perform the ***Absolute Zero Determination*** *demonstration. This demonstration is located in the Supplemental Lab Book.*

Strategies to Extend and Evaluate

- Have two groups of students debate whether or not the U.S. should convert to the metric system. Students on each team will try to convince a third group of students that the U.S. should or should not convert to the metric system. The rest of the students will evaluate the arguments and decide on a winning team by vote.
- Have students record what they think is the main idea of each section. Have pairs of students come to a consensus on each main idea. Then, have each pair combine with another pair and again come to a consensus. Finally, have each group share their results with the class. DI **Less Proficient Readers**

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 2.1 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. What are the basic units of measurement in the metric system for length and mass?
2. What unit is used to measure volume? How is this unit related to the measurement of length?
3. Explain the difference between *weight* and *mass.*
4. Give both the Celsius and Kelvin values for the boiling and freezing points of water.
5. How do you convert from Celsius to Kelvin? How does one degree Celsius compare with one Kelvin?
6. If someone told you that a swimming pool's temperature was 275 K, would it be safe for you to go for a swim?
7. Determine which metric measurement you would use for each of the following:
    a. The distance to the moon.
    b. The mass of a donut.
    c. The volume of a drinking glass.
    d. The length of your little finger.

2.2 Using Measurements

Key Concepts

In this lesson, students explore metric prefixes, scientific notation, and significant figures.

Lesson Objectives

- Use the metric system and its units.
- Convert between units.
- Use scientific notation in writing measurements and in calculations.
- Use significant figures in measurements. Unit conversions involve creating a conversion factor.

Lesson Vocabulary

scientific notation A shorthand way of writing very large or very small numbers. The notation consists of a decimal number between 1 and 10 multiplied by an integral power of 10. It is also known as exponential notation.

significant figures Any digit of a number that is known with certainty plus one uncertain digit. Beginning zeros and placeholder zeros are not significant figures.

Strategies to Engage

- Have students research odd measurement units such as the rood, fathom, or parasang. Facilitate a discussion with students about why the metric system is the measurement system used in chemistry.

Strategies to Explore

- Organize a metric scavenger hunt. Give each student a list of length, mass, and volume quantities expressed in metric units. Instruct students to look around the classroom and locate objects they think have those measurements. Instruct students to use rulers, balances, meter sticks, and graduated cylinders to measure those objects to see if their guesses were correct.
- Inform students that if they have difficulty determining whether or not a "0" in a measurement is significant, they can convert the measurement to scientific notation. If the 0 disappears, then it was not significant.

- Teach students the factor label method for conversions using the basic steps below. Have students practice using this method to perform metric conversions instead of simply moving the decimal point from left to right. This will prepare students to perform the more complex conversions they will need to be able to perform later on in this course.

1. Write the number and unit.
2. Set up a conversion factor.
   a. Place the given unit in the denominator.
   b. Place desired unit in the numerator
   c. Place a 1 in front of the larger unit.
   d. Determine the number of smaller units needed to make 1 of the larger unit.
3. Cancel units. Solve the problem.

Strategies to Extend and Evaluate

- Have students create a mnemonic device to help them memorize the metric prefixes.
- Have students write a short lesson that teaches other students the rules for determining the number of significant figures in a measurement. Instruct students to come up with examples for each rule.

Lesson Worksheets

Copy and distribute the worksheets in the *CK-12 Chemistry Workbook* named ***Significant Figures*** and ***Exponential Notation***. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 2.2 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Convert the following linear measurements:
   a. 0.01866 m = _______________ cm
   b. 2156 mm = _________________ m
   c. 15.38 km = ________________ m
   d. 1250.2 m = ________________ km
2. Convert the following mass measurements:
   a. 155.13 mg = _______________ kg
   b. 0.233 g = _________________ mg
   c. 1.669 kg = ________________ g
   d. 0.2885 g = ________________ mg
3. Write the following numbers in scientific notation:
   a. 0.0000479

b. $251,000,000$
c. 4260
d. 0.00206

4. How many significant figures are in the following numbers?
   a. 0.006258
   b. 1.00
   c. 1.01005
   d. 12500

2.3 Using Mathematics in Chemistry

Key Concepts

In this lesson, students explore the use of dimensional analysis and significant figures in chemistry problem solving.

Lesson Objectives

- Use units in problem solving.
- Do problem solving using dimensional analysis.
- Use significant figures in calculations.

Lesson Vocabulary

dimensional analysis A technique that involves the study of the dimensions (units) of physical quantities. It affords a convenient means of checking mathematical equations.

Strategies to Engage

- Write, "10 weeks" on the board and use dimensional analysis and unit conversions to quickly convert this quantity to seconds. Inform students that, in this lesson, they will learn to use dimensional analysis and unit conversions to perform complex conversions such as this.

Strategies to Explore

- Choose a place in the classroom to display the following rules for rounding to the correct number of significant figures in calculations. "When multiplying and dividing, limit and round to the least number of significant figures in any of the factors. When adding and subtracting, limit and round your answer to the least number of decimal places in any of the numbers that make up your answer." Have several students volunteer to write examples for each of these two rules.

Strategies to Extend and Evaluate

- Instruct students to begin with their age in years and use dimensional analysis and unit conversions to convert this value to hours and minutes. Have students express these values in scientific notation.

Lesson Worksheets

Copy and distribute the worksheet in the *CK-12 Chemistry Workbook* named ***Unit Conversions.*** Ask students to complete the worksheet alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 2.3 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

1. Perform the following calculations and give your answer with the correct number of significant figures:
   a. 0.1886×12
   b. $\frac{2.995}{0.16685}$
   c. $\frac{1210}{0.1223}$
   d. 910×0.18945
2. Perform the following calculations and give your answer with the correct number of significant figures:
   a. $10.5 + 11.62$
   b. $0.01223 + 1.01$
   c. $19.85 - 0.0113$
3. Do the following calculations *without a calculator:*
   a. $(2.0 \times 10^3)(3.0 \times 10^4)$
   b. $(5.0 \times 10^{-5})(5.0 \times 10^8)$
   c. $(6.0 \times 10^{-1})(7.0 \times 10^{-4})$
   d. $\frac{(3.0 \times 10^{-4})(2.0 \times 10^{-4})}{2.0 \times 10^{-6}}$

2.4 Using Algebra in Chemistry

Key Concepts

- In this lesson, students explore the use of algebra in chemistry problem solving.

Lesson Objectives

- Be able to rearrange mathematical formulas for a specific variable.
- Have an understanding of how to use units in formulas.
- Be able to express answers in significant figures and with units.

Lesson Vocabulary

None

Strategies to Engage

- When studying chemistry, students often ask, "When am I ever going to need this?" Inform students that many employers are looking to hire people with problem-solving skills. Facilitate a discussion with students about how solving chemistry problems gives them an opportunity to practice the problem-solving skills explored in algebra.

Strategies to Explore

- Use the mathematical calculations in this lesson to reduce the reliance on language skills. As you go through each example problem, use them to explain the concepts explored in this chapter. DI **English Language Learners**
- Perform the ***Density of Diet Soda vs. Regular Soda*** *demonstration. This demonstration is located in the Supplemental Lab Book.*
- Have students complete the lab ***Density Determination***. This lab is located in the Supplemental Lab Book.

Strategies to Extend and Evaluate

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 2.4 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. For the equation $\text{PV} = \text{nRT}$, re-write it so that it is in the form of " T =."
2. The equation for density is $\text{D} = \text{M}/\text{V}$. If D is $12.8\text{ g}/\text{cm}^3$, and M is 46.1 g, solve for V, keeping significant figures in mind.
3. The equation $P_1 V_1 = P_2 V_2$, known as Boyle's law, shows that gas pressure is inversely proportional to its volume. Re-write Boyle's law so it is in the form of $V_1 =?$.
4. The density of a certain solid is measured and found to be $12.68\text{ g}/\text{mL}$. Convert this measurement into kg/L.
5. In a nuclear chemistry experiment, an alpha particle is found to have a velocity of $14,285\text{ m}/\text{s}$. Convert this measurement into miles/hour.

Chapter 3 Chemistry in the Laboratory TE

Chapter Outline

Chemistry in the Laboratory

Outline

The chapter *Chemistry in the Laboratory* consists of four lessons that cover qualitative versus quantitative observation and data handling techniques.

- **Lesson 3.1 Making Observations**
- **Lesson 3.2 Making Measurements**
- **Lesson 3.3 Using Data**
- **Lesson 3.4 How Scientists Use Data**

Overview

In these lessons, students will explore:

- Qualitative and quantitative observations.
- The use of significant figures in measurements, accuracy and precision.
- Data patterns and graphs.
- Explore scientific laws, hypotheses and theories, and the construction of models in science.

Science Background Information

This information is provided for teachers who are just beginning to teach in this subject area.

- ***The Scientific Method and the Socratic Method***

The development of the scientific method was the result of centuries of cultural and societal evolution. Ranging from the philosophers of the Golden Age of Greece, through the applications of the Islamic scientists and into the ultimate

flowering of the Scientific Revolution. The main premise of the scientific method is the synthesis of a hypothesis and the collection of evidence, and the persistent application of experimentation designed to support or disprove that hypothesis.

Among the first practitioners of what developed into the scientific method was Al Hazen (965 – 1039), an Islamic mathematician renowned for his extensive studies in the fields of optics, physics and psychology. In particular, Al Hazen may have been among the very first to collect experimental evidence and to assemble his observations. For example, he conducted a series of tests on observing the light of external lanterns from an inner room to lead to the conclusion that the light emanated from the lanterns, not from the long held idea that light instead was the result of particles emerging from the eyes.

An alternative approach, called the Socratic method, consists of a method of inquiry in some ways following a parallel approach to the scientific method. The dialogues of Socrates, as collected by his student, Plato, consisted of framing a question, often about a philosophical dilemma, and addressing this issue with a logical answer. The strategy was pursued with series of questions intended to support or undermine the problem at hand. The goal of the Socratic method was to arrive at a conclusion via this sequence, mainly by uncovering any inconsistencies in their logic. This type of reasoning, utilizing only logic and the "thought experiment," lead to early misconceptions about the nature of physical realities. At times, the lack of simple experimentation produced erroneous conclusions that remained entrenched in many cultures for many years. The eminent philosopher Aristotle wrote about objects moving with "natural motion," that is, moving according to their composition and their speed, a result of their weight. More than a thousand years elapsed before the experiments conducted by Galileo rolling different objects down a ramp removed the role of weight in free fall acceleration.

Pacing the Lessons

Use the table below as a guide for the time required to teach the lessons of *Chemistry in the Laboratory.*

TABLE 3.1: Class Periods per Lesson

Lesson	Number of 60 Minute Class Periods
3.1 *Making Observations*	0.5
3.2 *Making Measurements*	0.5
3.3 *Using Data*	0.5
3.4 *How Scientists Use Data*	1.0

Managing Materials

The following items are needed to teach the strategies and activities described in the Teachers Edition of the FlexBook for *Chemistry in the Laboratory.*

TABLE 3.2: Chemistry in the Laboratory Materials List

Lesson	Strategy or Activity	Materials Needed
Lesson 3.1		
Lesson 3.2		
Lesson 3.3		
Lesson 3.4		

Multimedia Resources

You may find these additional web-based resources helpful when teaching *Chemistry in the Laboratory.*

- Graphing tutorial http://nces.ed.gov/nceskids/createagraph/default.aspx

Possible Misconceptions

Identify: Students may think that it is possible to measure a quantity with 100% accuracy.

Clarify: All measured values have some degree of uncertainty. Measurements are based on a comparison with a standard and can only be as accurate as the instrument that produced it.

Promote Understanding: Have students examine actual samples of each piece of equipment. Facilitate a discussion with students about the ability of each instrument to accurately measure 2.23 *mL* of water. Then, discuss with students the ability of the graduated pipet to accurately measure 2.23 *mL* of water. Explain to students that a degree of uncertainty is inherent in every measured value and that measurement instruments are not able to measure quantities with absolute accuracy.

Discuss: At the end of the lesson ask-Why is it not possible for a measured value to be absolutely accurate? (*All measured values have some degree of uncertainty.*)

Making the FlexBook Flexible

An important advantage of the FlexBook is the ability it gives you, the teacher, to select the chapters and lessons that you think are most important for your own classes. You should also consult the standards correlation table that follows when selecting chapters and lessons to include in the FlexBook for your classes.

TABLE 3.3: Standards Addressed by the Lessons in Chemistry in the Laboratory

Lesson	California Standards	SSES Standards	AAAS Standards
Lesson 3.1	1a		
Lesson 3.2	1a		
Lesson 3.3	1a		
Lesson 3.4	1a, 1e		

3.1 Making Observations

Key Concepts

In this lesson, students explore qualitative and quantitative observations.

Lesson Objectives

- Define qualitative and quantitative observations.
- Distinguish between qualitative and quantitative observations.
- Use quantitative observations in measurements.

Lesson Vocabulary

qualitative observations Describe the qualities of something and are described without numbers.

quantitative observations Observations that involve the use of numbers (quantities).

Strategies to Engage

- Show students an object such as a stapler or pencil sharpener. Ask students to describe the object. Facilitate a discussion with students about the types of observations that were made about the object.

Strategies to Explore

- Have students write a narrative of what they did in the morning from the time they woke up until the time they got to school. Facilitate a discussion about the qualitative and quantitative observations contained in the narratives.
- Perform the ***Separating Mixtures: Extracting Iron from Cereal*** *demonstration. This demonstration is located in the Supplemental Lab Book.*
- Have students complete the lab ***Chemical and Physical Changes***. This lab is located in the Supplemental Lab Book.

Strategies to Extend and Evaluate

- Have each student create a list of five observations. Instruct students to exchange papers with a classmate who will decide if each observation is qualitative or quantitative.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 3.1 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. The solutions are:

3.2 Making Measurements

Key Concepts

In this lesson students learn the use of significant figures in measurements, accuracy and precision.

Lesson Objectives

- Match equipment type, based on the units of measurements desired.
- Determine significant figures of the equipment pieces chosen.
- Define accuracy and precision.
- Distinguish between accuracy and precision.

Lesson Vocabulary

significant digits A way to describe the accuracy or precision of an instrument or measurement.

accuracy How close a number is to the actual or predicted value.

precision How close values are in an experiment to each other.

Strategies to Engage

- Students are likely to have heard about accuracy and precision in advertising and popular media. Call on volunteers to share with the class anything they may know about accuracy and precision. Point out correct responses, and clear up any misconceptions they have. Tell students they will learn more about accuracy and precision in this lesson.

Strategies to Explore

- Have students complete the lab ***Thermometer Calibration***. This lab is located in the Supplemental Lab Book.

Strategies to Extend and Evaluate

- Ask students to search for examples of the incorrect use of the terms "accuracy" and "precision" on the Web or in books. Have them quote the claim, reference the source, and then explain what is wrong.

Lesson Worksheets

There are no worksheets for this lesson

Review Questions

Have students answer the Lesson 3.2 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. A is neither accurate nor precise, B is both accurate and precise, C is precise but not accurate.
2. Their work is precise but not accurate.
3. The solutions are:
    1. 12
    2. 0.4
    3. 400
    4. 8.4
    5. 13
4. Answers may vary, but primarily because the units in the metric system are multiples of 10.
5. Accuracy deals with how close your values are to the true values, precision is how close the values are to each other.
6. The solutions are:
    1. 4
    2. 2
    3. 5
    4. 3
7. c
8. d

TABLE 3.4:

	Mass of Reactant 1	Mass of Reactant 2
Trial 1	1.45 ± 0.02 g	1.46 ± 0.02 g
Trial 2	1.43 ± 0.02 g	1.46 ± 0.02 g
Trial 3	1.46 ± 0.02 g	1.50 ± 0.02 g

9. The solutions are:
    1. 7.02

2. 400.00
3. 14
4. 4.8
5. 40 or 4×10^1

3.3 Using Data

Key Concepts

In this lesson students explore data patterns and graphs.

Lesson Objectives

- Recognize patterns in data from a table of values, pictures, charts and graphs.
- Make calculations using formulae for slope and other formulae from prior knowledge.
- Construct graphs for straight lines.
- Construct graphs for curves.
- Read graphs using the slope of the line or the tangent of the line.

Lesson Vocabulary

chemical reactivity An observation of the behavior of the element of compound based on its position in a reactivity (or activity) series.

periodic table An arrangement of elements in order of increasing atomic number.

alkali metals Group 1 metals of the periodic table (*H, Li, Na, K, Rb, Cs, Fr*).

alkaline earth metals Group 2 metals of the periodic table (*Be, Mg, Ca, Sr, Ba, Ra*).

density Measurement of a mass per unit volume. Density $= \frac{\text{mass}}{\text{volume}}$.

graphs Pictorial representation of patterns using a coordinate system ($x - y$ axis).

dependent variable The variable that changes depending on another variable ($y-$axis variable).

independent variable The variable that changes to cause another variable to change ($x-$axis variable).

Where the line crosses the $y-$axis.

conversion factor A ratio used to convert one unit to another.

linear relationship A relationship where the $x-$values change proportionally with the $y-$values leading to a straight line.

non-Linear relationship A relationship where the $x-$values do not change proportionally with the $y-$values leading to a curved line.

a line of best fit A line drawn on a scatter plot so that it joins as many points as possible and shows the general direction of the data. When constructing the line of best fit, it is also important to keep, approximately, an equal number of points above and below the line.

slope A formula to find the rate at which one factor is affecting the other.

$$m = \frac{y_2 - y_1}{x_2 - x_1}$$

tangent A straight line drawn to the curve.

solubility The amount of a substance that can dissolve in a given amount of solution.

Strategies to Engage

- Students are likely to be very familiar with the material explored in this section. Read each lesson objective and each statement in the lesson summary. Have students indicate their competency by using thumbs up or thumbs down to show "Yes" or "No." Whenever students use a thumbs down to show "No," use this as an opportunity to review the concept with the class.

Strategies to Explore

Strategies to Extend and Evaluate

- As a review of the lesson vocabulary, suggest that students make flash cards, with the vocabulary term on one side, and a drawing of what the term means on the other side. DI **English Language Learners**
- Have students write questions derived from Bloom's Taxonomy. Instruct students to research Bloom's taxonomy and write and answer one question from each of the six levels (knowledge, comprehension, application, analysis, synthesis, and evaluation.)

Lesson Worksheets

There are no worksheets for this lesson.

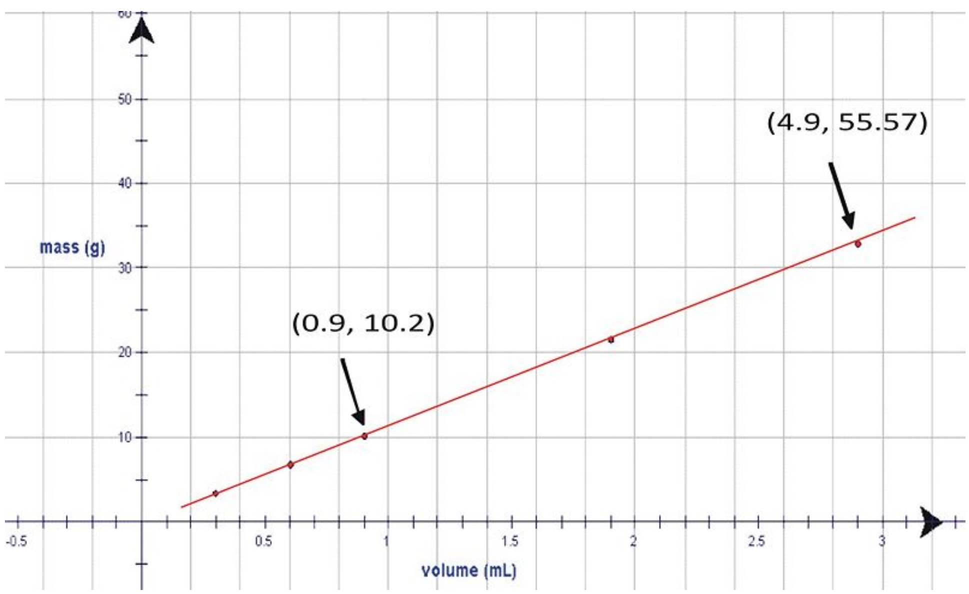
Review Questions

Have students answer the Lesson 3.3 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. the slope tells us the rate at which one factor is affecting the other
2. draw a tangent to the curve and find the slope of the tangent line
3. to change one unit into another
4. d
5. d
6. The solutions are:



1.

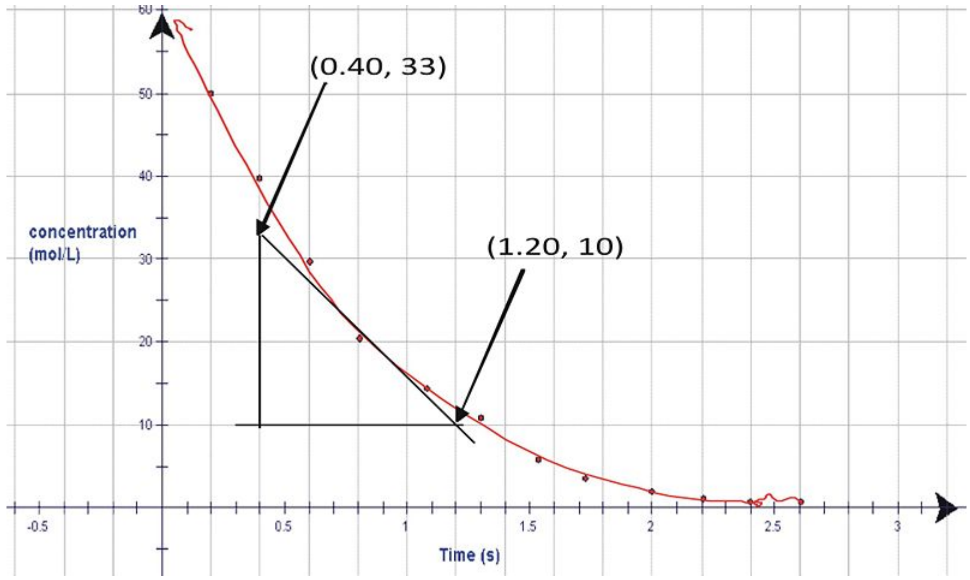
$$
\begin{aligned}
m &= \frac{y_2 - y_1}{x_2 - x_1} \\
m &= \frac{55.57 - 10.2}{4.9 - 0.9} \\
m &= 11.3 \text{g/mL}
\end{aligned}
$$

2. The slope represents the density of the unknown.
3. The unknown is lead.

TABLE 3.5:

Mass of Solid (g)	Volume of Solution (mL)
3.4	0.3
6.8	0.6
10.2	0.9
21.55	1.9
32.89	2.9
44.23	3.9
55.57	4.9

7. The solutions are:

1. 

2. The graph is a curve so draw the tangent and take the slope of the tangent:

$$m = \frac{y_2 - y_1}{x_2 - x_1} m \qquad = \frac{10 - 33}{1.20 - 0.40}$$

TABLE 3.6: Time (s) and Concentration

Time (s)	Concentration (mol/L)
0.20	49.92
0.40	39.80
0.60	29.67
0.81	20.43
1.08	14.39
1.30	10.84
1.53	5.86
2.00	1.95
2.21	1.07
2.40	0.71
2.60	0.71

8.

3.4 How Scientists Use Data

Key Concepts

In this lesson students explore scientific laws, hypotheses and theories, and the construction of models in science.

Lesson Objectives

- Define the terms law, hypothesis, and theory.
- Explain why scientists use models.

Lesson Vocabulary

natural laws A description of the patterns observed in the large amounts of data.

hypothesis An educated guess as to what is going to happen in the experiment.

theory Used to explain a law or to explain a series of facts/events.

law of conservation of mass Matter cannot be created nor destroyed.

model A description, graphic, or 3-D representation of theory used to help enhance understanding.

scientific method The method of deriving the theories from hypotheses and laws through experimentation and observation.

Strategies to Engage

- Give students examples of models in everyday life. For example, an ultrasound picture represents an unborn baby, a map represents an actual place, an athlete's list of statistics represents her performance. Facilitate a discussion with students about other examples of models in everyday life.

Strategies to Explore

- Have less proficient readers make a main ideas/details chart as they read the lesson. Instruct them to divide a sheet of paper down the middle and record the main ideas on the left side and the details for each main idea on the right side. Have students save their chart for reviewing lesson content. DI **Less Proficient Readers**

Strategies to Extend and Evaluate

- Encourage interested students to research science careers that use models. Students should be prepared to share their findings with the class.

Lesson Worksheets

There are no worksheets for this lesson.

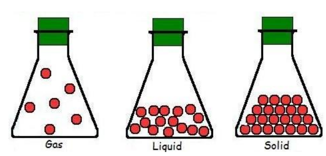
Review Questions

Have students answer the Lesson 3.4 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. (c)
2. (b)
3. The solutions are:



Chapter 4 Atomic Theory TE

Chapter Outline

Unit 2: Atomic Structure

Outline

This unit, Atomic Structure, includes four chapters that outlines the historical development of the atomic model and explains the structure of the atom.

- **Chapter 4 The Atomic Theory**
- **Chapter 5 The Bohr Model of the Atom**
- **Chapter 6 The Quantum Mechanical Model of the Atom**
- **Chapter 7 Electron Configurations for Atoms**

Overview

The Atomic Theory

The various models of the atom are developed from Dalton through Rutherford. This chapter also covers basic atomic structure and sub-atomic particles.

The Bohr Model of the Atom

This chapter introduces electromagnetic radiation, atomic spectra, and their roles in the development of the Bohr model of the atom.

Quantum Mechanics Model of the Atom

This chapter covers the quantum mechanical model of the atom, energy waves, standing waves, Heisenberg's uncertainty principle, and Schrodinger's equation. Quantum numbers, energy levels, energy sub-levels, and orbital shapes are introduced.

Electron Configurations for Atoms

This chapter covers electron spin, the Aufbau principle, and several methods for indicating electron configuration.

The Atomic Theory

Outline

The chapter *Atomic Theory* consists of three lessons in which the various models of the atom are developed from Dalton through Rutherford. This chapter also covers basic atomic structure and sub-atomic particles.

- **Lesson 4.1 Early Development of a Theory**
- **Lesson 4.2 Further Understanding of the Atom**
- **Lesson 4.3 Atomic Terminology**

Overview

In these lessons, students will explore:

- the development of atomic theory from the early Greek philosophers to Dalton's atomic theory.
- experiments leading to the discovery of subatomic particles and the development of atomic models.
- the structure of the atom.

Science Background Information

This background information is provided for teachers who are just beginning to teach in this subject area.

- Who Discovered the Neutron and How?

The construction of the modern atomic model consisting of the central nucleus and orbiting electrons was the result of years of experimentation and the dedication and insight of countless scientists. Yet, well into the twentieth century, the picture remained incomplete and inconsistencies and questions remained to be elucidated. By 1930, the fundamental positive and negative particles, the proton and the electrons had been identified and their dispositions relative to each other characterized. Ernest Rutherford, the discoverer of the proton, suggested the existence of what he termed a "proton-electron" pair, a heavy, yet neutral particle found in the nucleus, mainly on the basis of the differences between the atomic number (Z) of several atoms and their atomic mass. Further contemplation of another heavy fundamental particle arose to account for the unusual radiation emitted by beryllium atoms when bombarded by a stream of alpha particles. Irene and Frederic Joliot-Curie found this unusual radiation to be capable of ejecting protons, therefore had a mass comparable to that of protons. This was a confusing result, in that most physicists were under the assumption that this radiation better corresponded with the high energy but zero mass gamma radiation.

James Chadwick, who had worked for Ernest Rutherford at Manchester University and later at Cambridge University, replicated the Joliot-Curie experiment but with the intention of searching for a new fundamental neutral particle. Chadwick found that other light atoms other than beryllium gave off these new particles upon bombardment. He found the mass of this newly proposed particle to be about 10% greater than that of the proton, by comparing the velocity of the protons emitted by striking a hydrogen target with the neutral rays. He disproved the possibility of a proton-electron dual particle by illustrating that under no circumstances did the neutral ray particles degrade into smaller entities. For his efforts and insight, James Chadwick was awarded the Nobel Prize in Physics in 1935.

Since the discovery of the neutron, the advent of particle accelerators has produced evidence for hundreds of subatomic particles. Protons and neutrons are defined as baryons or heavy particles, whereas the list of leptons (light particles) is extensive.

Pacing the Lesson

Use the table below as a guide for the time required to teach the lessons of *Atomic Theory.*

TABLE 4.1: Class Periods per Lesson

Lesson	Number of 60 Minute Class Periods
4.1 Early Development of a Theory	1.0
4.2 Further Understanding of a Theory	1.0
4.3 Atomic Terminology	1.5

Managing Materials

The following materials are needed to teach the strategies and activities described in the Teachers Edition of the FlexBook for *Atomic Theory.*

TABLE 4.2: Atomic Theory Materials List

Lesson	Strategy or Activity	Materials Needed
Lesson 4.1		
Lesson 4.2	Lab Activity	shoe boxes, various objects, glue
Lesson 4.3	Exploration Activity	Copies of the periodic table

Multimedia Resources

You may find these additional Web-based resources helpful when teaching *Atomic Theory.*

- Jefferson Lab question archive http://education.jlab.org/qa/history_03.html
- Interactive atomic theory quiz http://www.visionlearning.com/library/quiz_taker.php?qid=6#38;mid=50
- Rutherford's gold foil experiment demonstration http://micro.magnet.fsu.edu/electromag/java/rutherford/

Possible Misconceptions

Identify: Students may think that the atom is larger than it really is.

Clarify: An atom is the smallest component of an element having the chemical properties of the element.

Promote Understanding: Inform students that a pure copper penny would contain about 2.4×10^{22} *atoms*. Have students write the number 2.4×10^{22} in standard form in order to see how large the number is. Explain to students that he population of the world is about 7 $billion(7 \times 10^9)$ people. So the number of copper atoms in a pure copper penny is more than three trillion times the population of the world.

Discuss: At the end of the lesson ask: "About how many atoms are there across the width of human hair?" (About a million.)

Making the FlexBook Flexible

An important advantage of the FlexBook is the ability it gives you, the teacher, to select the chapters and lessons that you think are most important for your own classes. You should also consult the standards correlation table that follows when selecting chapters and lessons to include in the FlexBook for your classes.

TABLE 4.3: Standards Addressed by the Lessons in Atomic Theory

Lesson	California Standards	SSES Standards	AAAS Standards
Lesson 4.1	1g, 1n		
Lesson 4.2	1e, 1h		
Lesson 4.3	1a, 1e		

4.1 Early Development of a Theory

Key Concepts

In this lesson, students will explore the development of atomic theory from the early Greek philosophers to Dalton's atomic theory.

Lesson Objectives

- State the Law of Definite Proportions.
- State the Law of Multiple Proportions.
- State Dalton's Atomic Theory, and explain its historical development.

Lesson Vocabulary

atomos (atomon) Democritus' word for the tiny, indivisible, solid objects that he believed made up all matter in the universe.

void Another word for empty space.

paradox Two statements that seem to be true, but contradict each other.

law of definite proportions In a given chemical substance, the elements are always combined in the same proportions by mass.

law of multiple proportions When two elements react to form more than one substance, the different masses of one element that are combined with the same mass of the other element are in a ratio of small whole numbers.

Strategies to Engage

- Introduce lesson concepts by facilitating a discussion about everyday observations.

Ask: What is wind? (Air movements.)

Ask: Have you ever seen wind? (No.)

Ask: How do you know that wind exists? (We can see its effects - tree limbs move, you can feel a temperature difference.)

Ask: What is an atom? (The smallest component of an element having the chemical properties of the element.)

Ask: Have you ever seen an atom with your own eyes? (No.)

Ask: What are some common observations that can be explained in terms of atoms? (Water evaporates, water erodes rocks, scents diffuse through a room.)

Explain to students that although they cannot see the atoms present in a material, there is plenty of evidence of their existence. Let them know that in this lesson, they will learn about the development of the idea of the atom.

Strategies to Explore

- Draw a line down the center of the board or chart paper. Write the law of definite proportions on one side, and the law of multiple proportions on the other side. Draw Figure 4.5 below the law of definite proportions. Instruct a student-volunteer to draw an example, similar to Figure 4.5, under the law of multiple proportions that illustrates the law. DI **English Language Learners**
- Write the basic assumptions of Dalton's atomic theory on the board or chart paper. Refer to it often. Facilitate discussions with students about the inaccuracies in Dalton's atomic theory as you explore the information in this chapter.
- Have students complete the lab ***Early Development of a Theory***. This lab is located in the Supplemental Lab Book.

Strategies to Extend and Evaluate

- Have students create a concept map relating the terms/objectives in the chapter.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 4.1 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. It turns out that a few of the ideas in Dalton's Atomic Theory aren't entirely correct. Are inaccurate theories an indication that science is a waste of time?
2. Suppose you are trying to decide whether to wear a sweater or a T-shirt. To make your decision, you phone two friends. The first friend says, "Wear a sweater, because I've already been outside today, and it's cold." The second friend, however, says, "Wear a T-shirt. It isn't logical to wear a sweater in July." Would you decide to go with your first friend, and wear a sweater, or with your second friend, and wear a T-shirt? Why?

3. Decide whether each of the following statements is true or false.
    a. Democritus believed that all matter was made of "atomos."
    b. Democritus also believed that there was only one kind of "atomos."
    c. Most early Greek scholars thought that the world was "ever-changing."
    d. If the early Greek philosophers hadn't been so interested in making gold, they probably would have liked the idea of the "atomos."
4. Match the person, or group of people, with their role in the development of chemistry.
    a. Early Greek philosophers - a. suggested that all matter was made up of tiny, indivisible objects
    b. alchemists - b. tried to apply logic to the world around them
    c. John Dalton - c. suggested that all matter was made up of tiny, indivisible objects
    d. Democritus - d. were primarily concerned with finding ways to turn common metals into gold
5. Early Greek philosophers felt that Democritus "atomos" theory was illogical because:
    a. no matter how hard they tried, they could never break matter into smaller pieces.
    b. it didn't help them to make gold.
    c. sulfur is yellow and carbon is black, so clearly "atomos" must be colored.
    d. empty space is illogical because it implies that nothing is actually something.
6. Which Law explains the following observation: carbon monoxide can be formed by reacting 12 grams of carbon with 16 grams of oxygen? To form carbon dioxide, however, 12 grams of carbon must react with 32 grams of oxygen.
7. Which Law explains the following observation: carbon monoxide can be formed by reacting 12 grams of carbon with 16 grams of oxygen? It can also be formed by reacting 24 grams of carbon with 32 grams of oxygen.
8. Which Law explains the following observation: 28 grams of carbon monoxide are formed when 12 grams of carbon reacts with 16 grams of oxygen?
9. Which Law explains the following observations: when 12 grams of carbon react with 4 grams of hydrogen, they produce methane, and there is no carbon or hydrogen left over at the end of the reaction? If, however, 11 grams of carbon react with 4 grams of hydrogen, there is hydrogen left over at the end of the reaction.
10. Which of the following is *not* part of Dalton's Atomic Theory?
    a. matter is made of tiny particles called atoms.
    b. during a chemical reaction, atoms are rearranged.
    c. during a nuclear reaction, atoms are split apart.
    d. all atoms of a specific element are the same.
11. Consider the following data: 3.6 grams of boron react with 1.0 grams of hydrogen to give 4.6 grams of BH_3. How many grams of boron would react with 2.0 grams of hydrogen?
12. Consider the following data: 12 grams of carbon and 4 grams of hydrogen react to give 16 grams of "compound *A*." 24 grams of carbon and 6 grams of hydrogen react to give 30 grams of "compound *B*." Are compound *A* and compound *B* the same? Why or why not?

4.2 Further Understanding of the Atom

Key Concepts

In this lesson, students explore the experiments leading to the discovery of subatomic particles and the development of atomic models.

Lesson Objectives

- Explain the experiment that led to Thomson's discovery of the electron.
- Describe Thomson's "plum pudding" model of the atom.
- Describe Rutherford's Gold Foil experiment, and explain how this experiment proved that the "plum pudding" model of the atom was incorrect.

Lesson Vocabulary

subatomic particles Particles that are smaller than the atom. The three main subatomic particles are electrons, protons and neutrons.

cathode rays rays of electricity that flow from the cathode to the anode. J.J. Thomson proved that these rays were actually negatively charged subatomic particles (or electrons).

cathode A negatively charged metal plate.

anode A positively charged metal plate.

cathode ray tube A glass tube with a cathode and anode, separated by some distance, at one end. Cathode ray tubes generate cathode rays.

phosphor A chemical that glows when it is hit by a cathode ray.

plum pudding model A model of the atom which suggested that the negative electrons were like plums scattered through the positive material (which formed the batter).

alpha Helium atoms that have lost their electrons. They are produced by uranium as it decays.

nucleus The small central core of the atom where most of the mass of the atom (and all of the atoms positive charge) is located.

Strategies to Engage

- Obtain about ten shoeboxes with lids. Glue a small object such as a toy, candle, or eating utensil inside of each shoebox. Place a marble into each shoebox. Instruct students to allow the marble to move around in the box and use the motion of the marble to guess the shape of the object inside of the shoebox. Explain to students this is similar to how they use clues to identify the object inside of the shoebox, scientists used many clues to discover the structure of the atom.

Strategies to Explore

- Draw a large circle on the board. As you explore this lesson, invite students to come to the board and change this "atomic model" to match how the model of the atom has changed over time. DI **English Language Learners**
- Have students complete the lab ***Rutherford's Experiment-Large Scale***. This lab is located in the Supplemental Lab Book.

Strategies to Extend and Evaluate

- Have students write a letter convincing the reader of the atom's existence and structure. Instruct students to include specific information about the experiments explored in this lesson.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 4.2 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Decide whether each of the following statements is true or false.
   a. Cathode rays are positively charged.
   b. Cathode rays are rays of light, and thus they have no mass.
   c. Cathode rays can be repelled by a negatively charged metal plate.
   d. J.J. Thomson is credited with the discovery of the electron.
   e. Phosphor is a material that glows when struck by cathode rays.
2. Match each observation with the correct conclusion.
   a. Cathode rays are attracted to a positively charged metal plate. - i. Cathode rays are positively charged. - ii. Cathode rays are negatively charged. - iii. Cathode rays have no charge.

b. Electrons have a negative charge. - i. atoms must be negatively charged. - ii. atoms must be positively charged. - iii. atoms must also contain positive subatomic material.
c. Alpha particles fired at a thin gold foil are occasionally scattered back in the direction that they came from - i. the positive material in an atom is spread throughout like the "batter" in pudding - ii. atoms contain neutrons - iii. the positive charge in an atom is concentrated in a small area at the center of the atom.

3. Alpha particles are:
    a. Helium atoms that have extra electrons.
    b. Hydrogen atoms that have extra electrons.
    c. Hydrogen atoms that have no electrons.
    d. Electrons.
    e. Helium atoms that have lost their electrons.
    f. Neutral helium atoms.
4. What is the name given to the tiny clump of positive material at the center of an atom?
5. Choose the correct statement.
    a. Ernest Rutherford discovered the atomic nucleus by performing experiments with aluminum foil.
    b. Ernest Rutherford discovered the atomic nucleus using a cathode ray tube.
    c. When alpha particles are fired at a thin gold foil, they never go through.
    d. Ernest Rutherford proved that the "plum pudding model" was incorrect.
    e. Ernest Rutherford experimented by firing cathode rays at gold foil.
6. Answer the following questions:
    a. Will the charges $+2$ and -2 cancel each other out?
    b. Will the charges $+2$ and -1 cancel each other out?
    c. Will the charges $+1$ and $+1$ cancel each other out?
    d. Will the charges -1 and $+3$ cancel each other out?
    e. Will the charges $+9$ and -9 cancel each other out?
7. Electrons are ______ negatively charged metals plates and ______ positively charged metal plates?
8. What was J. J. Thomson's name for electrons?
9. A "sodium cation" is a sodium atom that has lost one of its electrons. Would the charge on a sodium cation be positive, negative or neutral? Would sodium cations be attracted to a negative metal plate, or a positive metal plate? Would electrons be attracted to or repelled from sodium cations?
10. Suppose you have a cathode ray tube coated with phosphor so that you can see where on the tube the cathode ray hits by looking for the glowing spot. What will happen to the position of this glowing spot if:
    a. a negatively charged metal plate is placed above the cathode ray tube
    b. a negatively charged metal plate is placed to the right of the cathode ray tube
    c. a positively charged metal plate is placed to the right of the cathode ray tube
    d. a negatively charged metal plate is placed above the cathode ray tube, and a positively charged metal plate is placed to the left of the cathode ray tube
    e. a positively charged metal plate is placed below the cathode ray tube, and a positively charged metal plate is also placed to the left of the cathode ray tube.

4.3 Atomic Terminology

Key Concepts

In this lesson, students explore the structure of the atom.

Lesson Objectives

- Describe the properties of electrons, protons, and neutrons.
- Define and use an atom's atomic number (Z) and mass number (A).
- Define an isotope, and explain how isotopes affect an atom's mass, and an element's atomic mass.

Lesson Vocabulary

electron A type of subatomic particle with a negative charge.

proton A type of subatomic particle with a positive charge. Protons are found in the nucleus of an atom.

neutron A type of subatomic particle with no charge. Neutrons are found in the nucleus of an atom.

the strong nuclear force The force that holds protons and neutrons together in the nucleus of the atom. The strong nuclear force is strong enough to overcome the repulsion between protons.

atomic mass units (amu) A unit used to measure the masses of small quantities like protons, neutrons, electrons and atoms. It is useful, because the mass of a proton is very close to 1.0 *amu*.

elementary charge (The magnitude of charge on one electron or one proton. You can treat elementary charges as a unit of charge.

atomic number (An element's atomic number is equal to the number of protons in the nuclei of any of its atoms.

mass number (The mass number of an atom is the sum of the protons and neutrons in the atom.

Strategies to Engage

- Draw the following chart on the board or a sheet of chart paper.

TABLE 4.4:

	Proton	Neutron	Electron
Know			
Learned			

Ask students to think about what they know about the three subatomic particles explored so far in this chapter. Write some of their responses in the row marked "know". As you explore the information in this lesson, have student-volunteers write some of the information they have learned in the row marked "learned".

Strategies to Explore

- Give each student a copy of the periodic table. Call on students to tell you the atomic number and/or atomic mass given the name of the element and vice versa.
- You may want to compare the method used to calculate atomic mass from relative abundance to the method of calculating grades.

For example:

$$\begin{aligned}\text{Tests} &= (40\%)(0.78) = 31.2\\ \text{Labs} &= (20\%)(0.86) = 17.2\\ \text{Homework} &= (20\%)(0.90) = 18.0\\ \text{Quizzes} &= (20\%)(0.62) = 12.4\\ \text{Average} &= 78.8\end{aligned}$$

In this example, the percentages are analogs to the isotopic abundance. The average grade for each category would be the analog for mass number.

Strategies to Extend and Evaluate

- As a review of the lesson vocabulary, encourage students to make flash cards, with the vocabulary term on one side, and a definition or an example on the other side.
- Have students read the ***Timeline of the Discovery of the Elements***, and ***The Nature of Energy*** extra readings. These readings are located in the Supplemental Workbook.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 4.3 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Decide whether each of the following statements is true or false.
   a. The nucleus of an atom contains all of the protons in the atom.
   b. The nucleus of an atom contains all of the neutrons in the atom.
   c. The nucleus of an atom contains all of the electrons in the atom.
   d. Neutral atoms of a given element must contain the same number of neutrons.
   e. Neutral atoms of a given element must contain the same number of electrons.
2. Match the subatomic property with its description.
   a. electron - a. has an atomic charge of $+1e$
   b. neutron - b. has a mass of 9.109383×10^{-28} grams
   c. proton - c. is neither attracted to, nor repelled from charged objects
3. Arrange the electron, proton, and neutron in order of decreasing mass.
4. Decide whether each of the following statements is true or false.
   a. An element's atomic number is equal to the number of protons in the nuclei of any of its atoms.
   b. The symbol for an element's atomic number is (A).
   c. A neutral atom with $Z = 4$ must have 4 electrons.
   d. A neutral atom with $A = 4$ must have 4 electrons.
   e. An atom with 7 protons and 7 neutrons will have $A = 14$.
   f. An atom with 7 protons and 7 neutrons will have $Z = 14$.
   g. A neutral atom with 7 electrons and 7 neutrons will have $A = 14$.
5. Use the periodic table to find the symbol for the element with:
   a. 44 electrons in a neutral atom
   b. 30 protons
   c. $Z = 36$
   d. an atomic mass of 14.007 amu
6. When will the mass number (A) of an atom be...
   a. bigger than the atomic number (Z) of the atom?
   b. smaller than the atomic number (Z) of the atom?
   c. equal to the atomic number (Z) of the atom?
7. Column One contains data for 5 different elements. Column Two contains data for the same 5 elements, however different isotopes of those elements. Match the columns by connecting isotopes of the same element.

TABLE 4.5:

Column One	Column Two
a. an atom with 2 protons and 1 neutron	i. a C (carbon) atom with 6 neutrons
b. a Be (beryllium) atom with 5 neutrons	ii. an atom with 2 protons and 2 neutrons
c. an atom with $Z = 6$ and $A = 13$	iii. an atom with $Z = 7$ and $A = 15$
d. an atom with 1 proton and $A = 1$	iv. an atom with $A = 2$ and 1 neutron
e. an atom with $Z = 7$ and 7 neutrons	v. an atom with $Z = 4$ and 6 neutrons

Calculations:

8. Match the following isotopes with their respective mass numbers.
   a. an atom with $Z = 17$ and 18 neutrons - i. 35
   b. an H atom with no neutrons - ii. 4
   c. A He atom with 2 neutrons - iii. 1
   d. an atom with $Z = 11$ and 11 neutrons - iv. 23
   e. an atom with 11 neutrons and 12 protons - v. 22
9. Match the following isotopes with their respective atomic numbers.
   a. a B (boron) atom with $A = 10$ - i. 8
   b. an atom with $A = 10$ and 6 neutrons - ii. 2
   c. an atom with 3 protons and 3 neutrons - iii. 3
   d. an oxygen atom - iv. 4
   e. an atom with $A = 4$ and 2 neutrons - v. 5
10. Answer the following questions:
    a. What's the mass number of an atom that contains 13 protons and 13 neutrons?
    b. What's the mass number of an atom that contains 24 protons and 30 neutrons?
11. Answer the following questions:
    a. What's the mass number of the isotope of manganese (Mn) containing 28 neutrons?
    b. What's the mass number of the isotope of calcium (Ca) containing 20 neutrons?
12. Answer the following questions:
    a. What's the atomic number of an atom that has 30 neutrons, and a mass number of $A = 70$?
    b. What's the atomic number of an atom with 14 neutrons, if the mass number of the atom is $A = 28$?
13. Answer the following questions:
    a. What's the mass number of a neutral atom that contains 7 protons and 7 neutrons?
    b. What's the mass number of a neutral atom that contains 7 electrons and 7 neutrons?
    c. What's the mass number of a neutral atom that contains 5 protons, 5 electrons and 6 neutrons?
    d. What's the mass number of a neutral atom that contains 3 electrons and 4 neutrons
14. Answer the following questions:
    a. What element has 32 neutrons in an atom with mass number $A = 58$?
    b. What element has 10 neutrons in an atom with mass number $A = 19$?
15. Copper has two naturally occurring isotopes. 69.15% of copper atoms are $\text{Cu}-63$ and have a mass of 62.93 amu. The other 30.85% of copper atoms are $\text{Cu}-65$ and have a mass of 64.93 amu. What is the atomic mass of copper?

Chapter 4 Enrichment

Extra Readings

The Nature of Energy

The Four Fundamental Forces

There are four fundamental forces within all atoms, that dictate interactions between individual particles, and the large-scale behavior of all matter throughout the universe. They are the strong nuclear force, the weak nuclear force, the electromagnetic force, and the gravitational force.

Gravity is a force of attraction that acts between each and every particle in the universe. It is always attractive, never repulsive. It pulls matter together. It is gravity that keeps the planets in their orbits around the sun, the moon in its orbit around the earth, binds galaxies together in clusters, causes apples to fall from trees, and keeps you standing on the earth.

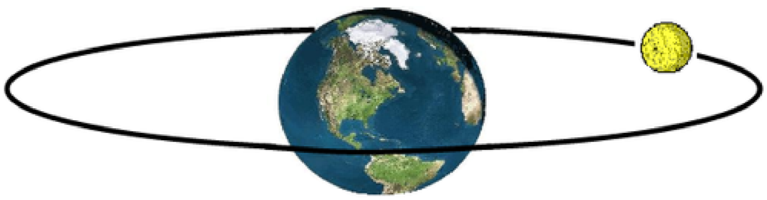


FIGURE 4.1

The electromagnetic force determines the ways in which electrically charged particles interact with each other and also with magnetic fields. This force can be attractive or repulsive. Like charges (two positive or two negative charges) repel each other; unlike charges attract. The electromagnetic force binds electrons in electron clouds around the positively charged nucleus of an atom and also governs the emission and absorption of light and other forms of electromagnetic radiation. Since the outside of atoms is an electron cloud, the electromagnetic force controls the interaction of materials when they touch each other and thus is the cause of the existence of liquids and solids and allows you to talk, move, breathe, and so on. All of the interactions between objects that you see every day is controlled by the electromagnetic force.

The strong nuclear force is the force that binds the atomic nucleus together. You may not have thought about it at the time the atomic nucleus was introduced to you but the atomic nucleus contains a number of positively charged protons held tightly together in a tiny space. From what we know about the repulsive force between like charges, the atomic nucleus should not stay together. The positive protons should repel each other strongly and fly apart. The fact that the protons and neutrons stay together in an atomic nucleus is because they are held there by an extremely strong force – namely, the strong nuclear force. Both the strong and weak nuclear forces operate only when the particles being attracted are extremely close together.

At this level, the weak nuclear force will skipped over with little consideration. We will just note that protons and neutrons are also composed of smaller particles (quarks, etc.) and these particles have a force which holds them together to form protons and neutrons. This force is the weak nuclear force.

Energy

Energy, like matter, is an important factor in our universe. Without energy, all matter – living and non-living – would be at a standstill; nothing would move, nothing would live. Energy is considered to be the "mover of matter". The idea of energy is one that unites all the sciences. Energy does not have mass and does not take up space so it is not matter. Energy is defined as the ability to do work. An example of doing work (in the physics sense) is when you lift an object from the ground onto a table. The amount of work done depends on the force you had to apply to lift the object (its weight – or – the force of gravity on it) and the height you lifted the object. The greater the weight of the object or the higher it is lifted, the greater the amount of work done.

Energy comes in many forms. Besides mechanical energy, there is heat, light, sound, electricity, magnetism, chemical, and nuclear energy. Almost any form of energy can be converted into any other form. Our chief source of energy is the sun. It provides us with light, which can then be converted into other forms of energy. Light can be absorbed by matter and converted into heat. Light can also be absorbed by plants in the process of photosynthesis and be converted into chemical energy.

Kinetic and Potential Energy (Mechanical Energy)

Energy can be classified as either kinetic energy or potential energy. The original definition we gave for energy was the ability to do work. The ability to do work could also be stated as the "ability to make matter move". Anything that can make a piece of matter move has energy. It should be obvious that a moving object has the ability to make another piece of matter move simply by colliding with it. Therefore, all moving objects have energy. This type of energy (the energy of moving objects) is called kinetic energy.

There are also non-moving objects that have the ability to make matter move. These objects have the ability to make matter move because of their position. For example, a rock held up in the air has the ability to make matter move – all that is required is that whatever is holding the rock up must release it. A stretched rubber band has the ability to make matter move – all that is required is that whatever is holding the rubber band must release it. This type of energy is stored energy or potential energy.

This baseball flying through the air has both kinetic and potential energy. The kinetic energy is due to its motion and the potential energy is due to the balls' height above the ground.

The kinetic energy of an object can be calculated by multiplying one-half of its mass times its velocity squared.

$$KE = \frac{1}{2}mv^2$$

The gravitational potential energy of an object can be calculated by multiplying the mass of the object times the acceleration due to gravity times the height the object can fall.

$$PE = mgh$$

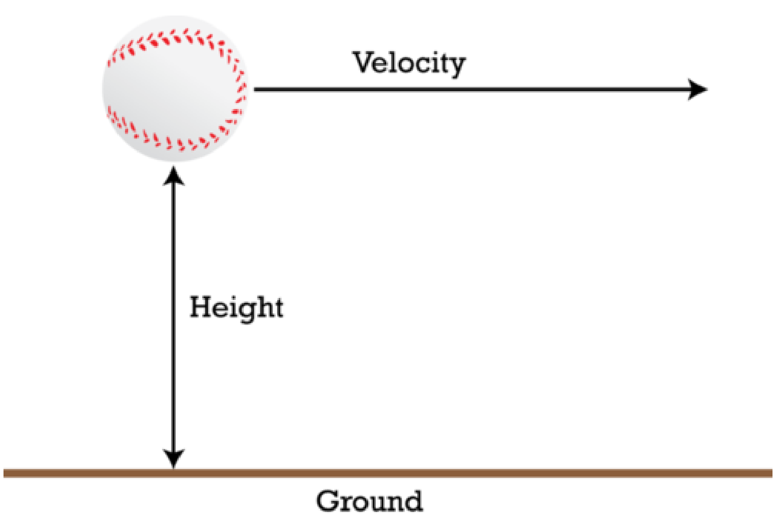


FIGURE 4.2

Potential energy is always present when two objects are attracted or repelled and are held in position. The most obvious case is an object that is held above the earth. The object is attracted to the earth by gravity but is kept from falling (gravitational potential energy). This same type of energy is present in bent sticks, compressed or stretched springs, stretched rubber bands, like or unlike electrical charges, and like or unlike magnetic poles. In all these cases, the potential energy can be calculated by multiplying the force of attraction (or repulsion) by the distance one object will move.

Energy Transmission

Scientists use three words to indicate the different methods by which energy moves from place to place. These three words are conduction, convection, and radiation.

Conduction

We are all familiar with the concept of molecules in constant random motion. This molecular motion increases when we heat the molecules and decreases when we cool the molecules. The energy of these moving molecules is kinetic energy. Kinetic energy is transferred between molecules when molecules at different temperatures collide with each other. When molecules at different temperatures collide with each other, energy is transferred from the "hotter" molecules to the "colder" molecules. Consider an object such as an iron bar – we can view the bar as a long chain of molecules crowded very close together. Remember that molecules or atoms in a solid are in a tightly packed pattern.

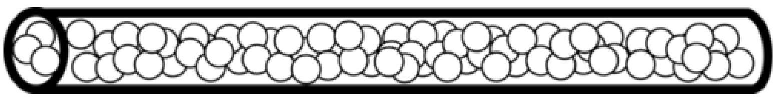


FIGURE 4.3

If this bar lies on a table for a few minutes, all the particles will be at about the same temperature. This is because each molecule is constantly bumping into its neighbors and these collisions transfer kinetic energy from a faster moving particle to a slower moving particle. If one end of this bar is placed into a flame and heated, the bar particles that are in the flame will get very hot.

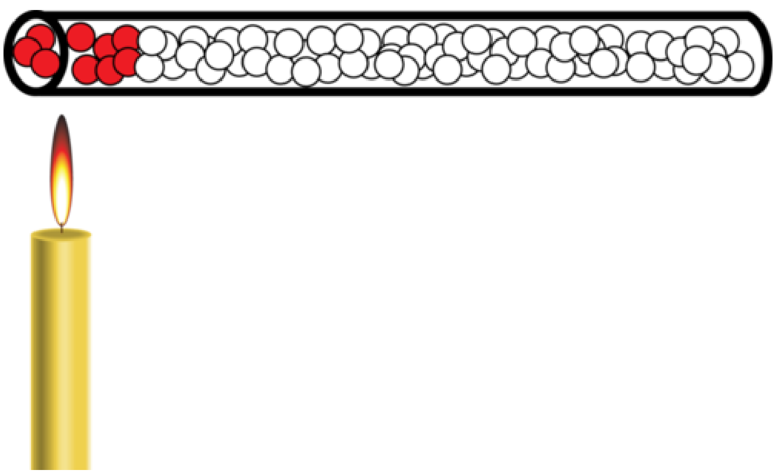


FIGURE 4.4

When the hot particles bump into cold ones, the cold particles gain kinetic energy from the hot ones and thus the cold ones also become hot. Those particles then bump into their neighbors down the bar and eventually, all the molecules in the bar will be hot. This process of passing heat (kinetic energy) from particle to particle by collision is known as conduction.

Any time two objects at different temperatures touch each other, heat will be conducted from the hot one to the cold one by this process.

Molecules that make up living systems (like you) are called organic molecules. Most of these molecules are much more fragile than non-organic molecules. Non-organic molecules can usually reach quite a high temperature before the molecules are damaged. Organic molecules, however, are frequently long chains of carbon atoms and are easily to break if they are jerked around. If a hot object conducts heat to your hand, like all other conduction, the increased

temperature causes the molecules of your hand to move around more rapidly. At temperatures at or below $40°C$, your nervous system reacts in such a way that the sensation is not unpleasant. But at higher temperatures, your molecules begin moving around so rapidly that some of them break apart. When this happens, your nervous system sends a signal to your brain that you are in PAIN so that you will remove your hand from the heat as fast as possible. If many molecules are broken, the tissue is permanently damaged and must be replaced (healed) by the body.

FIGURE 4.5

Convection

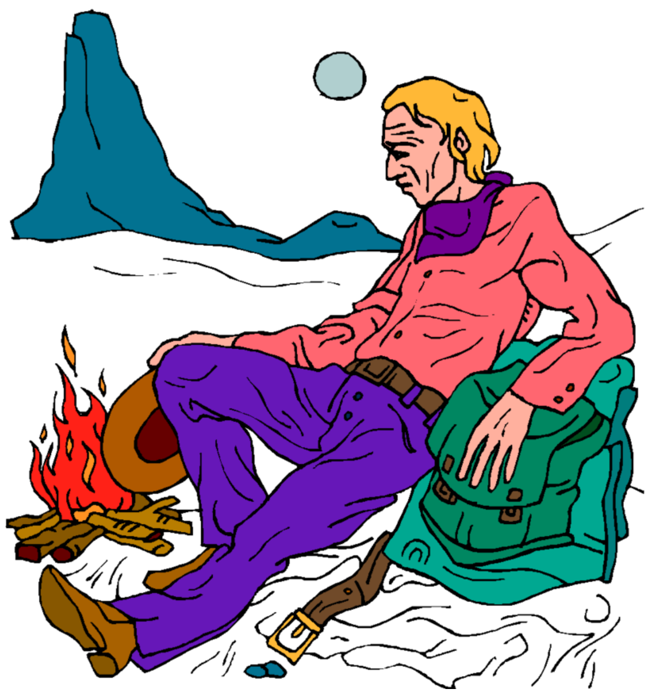
Another way to move heat (energy) from one place to another is to heat up some substance, like air or water for example, and then to move the heated substance to another place. Essentially, the matter holds the energy in the form of heat, and when you move the matter to another place, you are also moving the energy it contains. In most homes, we use a furnace to heat air and use a fan to blow the hot air through ducts to various places in the house to warm it. In some places, water is heated and then pumped through pipes and into radiators to transport the heat from where it is produced to other areas. The process of moving matter that contains heat to other places is called convection.

Nature has its own convection system of heating up air in one place and then wind blows the air to another location. There are also convection currents in lakes and oceans, where cold water sinks and warm water rises causing water

flow (and thus heat transfer) from place to place.

Radiation

When you build a campfire, the heat produced from the flames mostly goes upward. This is because hot air is less dense than cold air and so the hot air behaves like a helium balloon and goes straight up. If you stand a few feet to the side of a campfire, however, you will also feel heat coming from the fire. This heat does not get there by conduction or convection. This heat arrives at your position by radiation. Radiation is a type of energy transfer that can occur even through a vacuum – it needs no air or any other matter to carry it. This is quite different from conduction and convection which both require molecules of matter to transfer the energy.



This is the way that light from the sun travels through the vacuum of outer space and arrives at the earth. This type of energy is called electromagnetic radiation. There are various levels of energy for electromagnetic radiation. One of the lower energy forms of EMR are radio waves. As the energy of EMR increases, we encounter infra-red light, visible light, ultra-violet light, and x-rays. The highest energy form of EMR is gamma radiation. The radio and television signals we use for communication are electromagnetic radiation. Astronauts in outer space can still communicate with people on earth because the radio and TV signals do not need matter to travel through – they can travel through a vacuum.

Infra-red light is frequently used in remote control devices like your TV remote. The energy that cooks your food in a microwave oven is EMR. Radio signals are used to operate automatic garage door openers and infra-red "eyes" are used to stop the garage door if something is in the way of the door. Doctors and dentists use x-rays to make pictures of bones and teeth. X-rays are such a powerful form of EMR that the energy passes through skin, flesh, and many other substances. Electromagnetic radiation at the level of x-rays and gamma rays are so powerful that they can be dangerous to human beings.

Ultraviolet light is another form of EMR. UV light is the part of the sunlight responsible for tanning skin, burning skin, and in some cases, causing skin cancer.

For humans, the most common form of EMR is visible light, which our eyes use for vision.

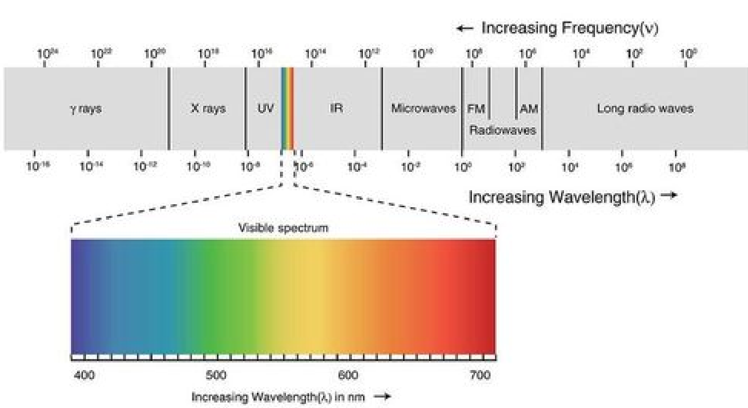


FIGURE 4.6

A Little More on Gravity

Every particle of matter attracts every other particle of matter with a force which Isaac Newton named the force of gravity. We know this force exists, we can calculate the size and direction of the force, but we cannot yet explain how or why the force works. This force between particles of matter exists everywhere in the universe and it attracts all matter everywhere in the same way.

The size of the force of gravity is dependent on the masses of the two attracting objects and also on the distance between the centers of the two objects. The amount of matter in an object is called its mass and whenever the mass of one or both of the objects is increases, the force of gravity between the objects also increases. If the objects are brought closer together, the force of gravity increases and if the objects are moved farther apart, the force of gravity decreases.

The force of gravity between two small objects, such as two people, who are standing one meter apart is so small that we cannot even measure it. When one of the objects is very large, such as the earth, the force of gravity becomes large. The force of gravity pulling on a 50 *kilogram* person standing on the surface of the earth would be about 500 *Newtons*. In this English system, this force would correspond to about 110 *pounds*. The weight of a person is, in fact, the force of gravity acting on that person. If you weigh 120 *pounds*, it means that the earth is pulling on you with a force of 120 *pounds*.

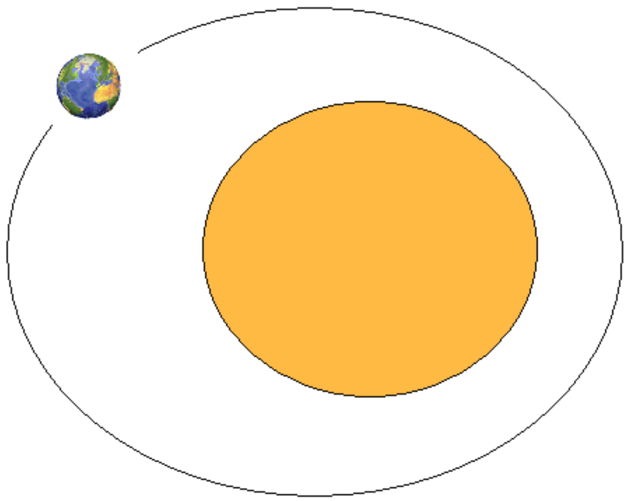
FIGURE 4.7

If two objects the size of the earth were right next to each other, the force of gravity between them would be about 10,000,000,000,000,000,000,000,000 *Newtons*. As you can see, small objects have very small attractions to due

gravity but very large objects can produce a gigantic force of gravity.

If we hold two $1,000,000\ kg$ objects (like a battleship) one meter apart, the force of gravity between them would be about $70\ Newtons$. If we move the objects to a $10\ meter$ separation, the force would become about $0.7\ N$ and if we separate the objects by $100\ meters$, the force becomes $0.007\ Newton$. The force of gravity weakens rapidly as the distance between objects increases.

With two really large objects, like the earth and the sun, there is a certain distance where the force of gravity is just enough to keep the objects circling each other. The smaller object does most of the moving in this circle (actually, more like an oval).



Timeline for the Discovery of the Elements

TABLE 4.6: **Discovery of the Elements by Year**

Method of Discovery	Year of Discovery	Element Symbols
Found Free in Nature or Simple Metallurgy	Before 1 A.D.	C, S, Hg, Sn, Pb, Fe, Cu, Ag, Au
Simple Metallurgy	Alchemists to 1735	Zn, P, Bi, Sb, As
Simple Metallurgy	1735 – 1745	Pt, Co
Simple Metallurgy	1745 – 1755	Ni
	1755 – 1765	
Electrochemistry	1765 – 1775	F, Mn, Cl, O, N, H
Electrochemistry	1775 – 1785	Te, W, Mo
Electrochemistry	1785 – 1795	Y, Ti, Sr, U
Electrochemistry	1795 – 1805	Ir, Os, Rh, Pd, Ce, Ta, Nb, Be, Fr, V
Electrochemistry	1805 – 1815	I, B, Mg, Ca, Ba, K, Na
Electrochemistry	1815 – 1825	Zr, Si, Se, Cd, Li
Electrochemistry	1825 – 1835	Th, Br, Al
Electrochemistry	1835 – 1845	Ru, Er, Tb, La
	1845 – 1855	
Electrochemistry	1855 – 1865	In, Tl, Rb, Cs
	1865 – 1875	
Separation Techniques	1875 – 1885	Tm, Ho, Sc, Sm, Yb, Ga
Separation Techniques	1885 – 1895	Ar, Ge, Dy, Gd, Nd, Pr

TABLE 4.6: (continued)

Method of Discovery	Year of Discovery	Element Symbols
Separation Techniques	1895 – 1905	Rn, Ac, Ra, Po, Xe, Ne, Kr, Eu, He
Separation Techniques	1905 – 1915	Lu
Separation Techniques	1915 – 1925	Pa, Hf
Separation Techniques	1925 – 1935	Re
Nuclear Synthesis	1935 – 1945	Cm, Pu, Np, At, Fr, Tc
Nuclear Synthesis	1945 – 1955	Cf, Bk, Pm, Am, Es, Fm, Md
Nuclear Synthesis	1955 – 1965	No, Lr, Rf, Db, Sg, Bh, Hs, Mt
Nuclear Synthesis	1965 – Present	Ds, Rg, Cp, 114, 116, 118

4.4 References

1. Richard Parsons. . CCBYSA
2. Richard Parsons. . CCBYSA
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6. Richard Parsons. . CCBYSA
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CHAPTER 5 The Bohr Model TE

Chapter Outline

The Bohr Model of the Atom

Outline

The chapter *The Bohr Model of the Atom* consists of four lessons that introduce electromagnetic radiation, atomic spectra, and their roles in the development of the Bohr model of the atom.

- **Lesson 5.1 The Wave Form of Light**
- **Lesson 5.2 The Dual Nature of Light**
- **Lesson 5.3 Light and the Atomic Spectra**
- **Lesson 5.4 The Bohr Model**

Overview

In these lessons, students will explore:

- The wave form model of light.
- The experiments that led to the concept of wave-particle duality.
- Continuous and discontinuous spectra.
- The explanations provided by the Bohr atomic model as well as its limitations.

Science Background Information

This background information is provided for teachers who are just beginning to teach in this subject area.

- What is the Electromagnetic Spectrum?

The visible light or radiant energy that illuminates our portion of the universe and enriches our existence with the appearance of different colors and hue intensities, was the first type of electromagnetic radiation evident to mankind.

The remaining regions of the electromagnetic spectrum have only recently been elucidated. These varied regions can be differentiated on the basis of their wavelength (in length units of meters or millimeters), or frequency (in sec^{-1} or Hertz units).

The first region other than the section of the spectrum visible to human eyes was the infrared portion. William Herschel, also known as the discoverer of the first planet to be revealed in modern times, Uranus, was responsible for slowing rays of light with a prism, and redirecting the light rays into heat-absorbing bulbs. He found that the "caloric rays" were most intense beyond the red portion, producing the highest absorption temperatures yet the rays could be refracted and reflected like visible light.

In Germany, Johann Ritter, learning about Herschel's discovery, attempted to identify the complementary radiation beyond the violet region of the visible spectrum by exposing silver chloride crystals to refracted sunlight. Ritter originally called this new discovery "chemical radiation" but in time, this radiation became known as ultraviolet (beyond the violet).

James Clerk Maxwell created the Electromagnetic Theory, which served to unify the initially disparate fields of electricity and magnetism utilizing Maxwell's Equations. His work suggested that light itself was one of several types of electromagnetic waves, all traveling at the velocity of light, c.

The next portion of the electromagnetic spectrum to be identified was located in the low energy region. In 1887, German physicist Heinrich Hertz added very long wavelength radio waves to the spectrum, but his research did not pursue applications of this technology as he felt that there was no practical use for it. It was left to Nicola Tesla and Guglielmo Marconi to find ways to utilize "wireless telegraphy" for the public.

The discovery of X-rays followed soon thereafter. Wilhelm Roentgen, a Bavarian physicist, studied the passage of cathode rays from an induction coil through a glass tube that had been partially evacuated. He noticed that these rays when projected upon a fluorescent screen caused it to glow. Roentgen also found that these rays penetrated skin and could cast an image of the bones within upon on photographic plate. The first X-ray image published was that of Frau Roentgen's hand.

Interest in uncovering new elements and new phenomena such as X-rays was all consuming as the end of the nineteenth century approached. Henri Becquerel, in Paris, discovered that uranium salts were the source of radioactivity. Another Parisian researcher, Paul Villard, also studied radioactive sources and in 1900, established that certain radioactive materials emitted what become known as gamma rays, high-energy radiation with even shorter wavelengths than X-rays.

By the early twentieth century, most of the regions of the electromagnetic spectrum had been explored and applications of the different manifestations such as radio waves and X-rays had been explored. One region, however remained largely unexamined until the 1940s. This type of electromagnetic radiation, initially known as ultrashort radio waves, consisting of wavelengths in the $1\ meter$ – one millimeter range, was used to send radar signals to establish distance. Use of this application expanded during World War II. Engineers at the Raytheon corporation building the vacuum tubes for military uses noticed that the heat emitted by the tubes could be used to warm their hands in the winter months. The idea of incorporating this technology to construct microwave ovens was implemented by Raytheon engineers John Spencer and Marvin Bock. The ubiquitous modern cell phones also utilize microwave radiation to send signals, at an intensity level too low to result in thermal heating.

New applications are continually being added to the complement of uses for the different ranges of wavelengths and frequencies encompassed by the electromagnetic radiation, shedding "light" on previously unexplored areas of potential technology.

Pacing the Lessons

Use the table below as a guide for the time required to teach the lessons of *The Bohr Model of the Atom.*

TABLE 5.1: Class Periods per Lesson

Lesson	60 Minute Class Periods per Lesson
5.1 *The Wave Form of Light*	1.5
5.2 *The Dual Nature of Light*	1.0
5.3 *Light and the Atomic Spectra*	1.0
5.4 *The Bohr Model*	1.0

Managing Materials

The following items are needed to teach the strategies and activities described in the Teachers Edition of the FlexBook for *The Bohr Model of the Atom.*

TABLE 5.2: The Bohr Model of the Atom Materials List

Lesson	Strategy or Activity	Materials Needed
5.1	Exploration Activity	metric ruler
5.2	Exploration Activity	salt
5.3	Engagement Activity	wintergreen mints, pliers, clear tape
5.4		

Multimedia Resources

You may find these additional Web-based resources helpful when teaching *The Bohr Model of the Atom.*

- http://www.learner.org/teacherslab/science/light/ Interactive lesson on light and color.
- http://www.teachersdomain.org/resource/phy03.sci.phys.energy.nasaspectrum/ Electromagnetic spectrum video.

Possible Misconceptions

Identify: Students may think exposure to any amount of electromagnetic radiation is harmful.

Clarify: Electromagnetic radiation is the transfer of energy in the form of electromagnetic waves. All of the parts of the electromagnetic spectrum are referred to as electromagnetic radiation. We are constantly bombarded by natural sources of radiation. Although the highest frequency electromagnetic waves can be beneficial, absorbing too much of these forms of radiation can be harmful to the body.

Promote Understanding: Have students research various applications of electromagnetic radiation.

Discuss: At the end of the lesson ask "Under what circumstance is electromagnetic radiation harmful?". (When the body absorbs too much.)

Making the FlexBook Flexible

An important advantage of the FlexBook is the ability it gives you, the teacher, to select the chapters and lessons that you think are most important for your own classes. You should also consult the standards correlation table that follows when selecting chapters and lessons to include in the FlexBook for your classes.

TABLE 5.3: Standards Addressed by the Lessons in The Bohr Model of the Atom

Lesson	[California] Standards	[NSES] Standards	[AAAS] Benchmarks
Lesson 5.1			
Lesson 5.2	1h, 1j		
Lesson 5.3			
Lesson 5.4	1i, 1j		

5.1 The Wave Form of Light

Key Concepts

In this lesson, students explore the waveform model of light.

Lesson Objectives

- The student will define the terms wavelength and frequency with respect to waveform energy.
- The student will state the relationship between wavelength and frequency with respect to electromagnetic radiation.
- The student will state the respective relationship between wavelengths and frequencies of selected colors on the electromagnetic spectrum.

Lesson Vocabulary

crest High point in a wave pattern (hill).

trough Low point in a wave pattern (valley).

amplitude of a wave The 'height' of a wave. In light waves, the amplitude is proportional to the brightness of the wave.

frequency of a wave (v) The 'number' of waves passing a specific reference point per unit time. The frequency of a light wave determines the color of the light.

hertz (Hz) The SI unit used to measure frequency. One Hertz is equivalent to 1 event (or one full wave passing by) per second.

wavelength The length of a single wave from peak to peak (or trough to trough). The wavelength of a light wave determines the color of the light.

electromagnetic spectrum A list of all the possible types of light in order of decreasing frequency, or increasing wavelength, or decreasing energy. The electromagnetic spectrum includes gamma rays, X-rays, UV rays, visible light, IR radiation, microwaves and radio waves.

Strategies to Engage

- Students are likely to have heard about various forms of electromagnetic radiation in popular media (e.g.,gamma rays, x-rays, microwaves). Call on volunteers to share with the class anything they already know about electromagnetic waves. Point out correct responses and clear up any misconceptions. Tell students they will learn more about electromagnetic waves in this lesson.

Strategies to Explore

- Have each student draw a waveform similar to Figure 5.1. Instruct students to label one wavelength, a crest, and a trough. Have students use a ruler to measure the wavelength to the nearest tenth of a centimeter. Divide students into groups of three or four. Instruct students to compare their sketches and measurements and rank their waveforms from highest to lowest frequency. (The higher the frequency the longer the wavelength). DI **English Language Learners**
- Point out to students that for the most part the term "light" is used to describe the visible portion of the electromagnetic spectrum.

Strategies to Extend and Evaluate

- Ask students to search for examples of myths regarding electromagnetic radiation on the Web or in books. Have them quote the claim, reference the source, and then explain what is wrong.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 5.1 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Choose the correct word in for the following statement. Blue light has a (*longer* or *shorter*) wavelength than red light.

2. Choose the correct word in for the following statement. Yellow light has a (*higher* or *lower*) frequency than blue light.

3. Choose the correct word in for the following statement. Green light has a (*larger* or *smaller*) energy than red light.

4. If "light A" has a longer wavelength than "light B", then "light A" has ________________ "light B".
   A. a lower frequency than B. a higher frequency than C. the same frequency as

5. If "light C" has a shorter wavelength than "light D", then "light C" has ________________ "light D".
   A. a larger energy than B. a smaller energy than C. the same energy as

6. If "light E" has a higher frequency than "light F", then "light E" has ___________________ "light F".
   A. a longer wavelength than B. a shorter wavelength than C. the same wavelength as

7. If "light G" has a higher frequency than "light H", then "light G" has ___________________ "light H".
   A. a larger energy than B. a smaller energy than C. the same energy as

8. If "light J" has larger energy than "light K", then "light J" has ___________________ "light K".
   A. a shorter wavelength than B. a longer wavelength than C. the same wavelength as

9. Which of the following statements is true?
   A. The frequency of green light is higher than the frequency of blue light and the wavelength of green light is longer than the wavelength of blue light. B. The frequency of green light is higher than the frequency of blue light and the wavelength of green light is shorter than the wavelength of blue light. C. The frequency of green light is lower than the frequency of blue light and the wavelength of green light is shorter than the wavelength of blue light. D. The frequency of green light is lower than the frequency of blue light and the wavelength of green light is longer than the wavelength of blue light. E. The frequency of green light is the same as the frequency of blue light and the wavelength of green light is shorter than the wavelength of blue light.

10. As the wavelength of electromagnetic radiation increases:
    A. its energy increases. B. its frequency increases. C. its speed increases. D. more than one of the above statements is true. E. none of the above statements is true.

11. List three examples of electromagnetic waves.
12. Why do white objects appear white?
13. Name the colors present in white light in order of increasing frequency.
14. Why do objects appear black?

5.2 The Dual Nature of Light

Key Concepts

In this lesson students explore the experiments that led to the concept of wave-particle duality.

Lesson Objectives

- Explain the double-slit experiment and the photoelectric effect.
- Explain why light is both a particle and a wave.
- Use and understand the formula relating a light's velocity, frequency, and wavelength, $c = f\lambda$.
- Use and understand the formula relating a light's frequency and energy, $E = hf$.

Lesson Vocabulary

diffraction The tendency of a wave to spread out in a circular shape when passed through a small opening.

double-slit experiment When light is passed through two narrowly separated openings (slits), the light produces a resulting pattern of peaks and troughs that suggests that light behaves like a wave.

photoelectric effect The process whereby light shone on a metal surface knocks electrons (called photoelectrons) off of the surface of the metal.

black-body radiation Light produced by a black object when the object is heated.

photon or quanta of light A tiny particle-like packet of energy.

wave-particle duality of light Einstein's theory, which concluded that light exhibits both particle and wave properties.

electromagnetic spectrum A list of all the possible types of light in order of decreasing frequency, or increasing wavelength, or decreasing energy. The electromagnetic spectrum includes gamma rays, X-rays, UV rays, visible light, IR radiation, microwaves and radio waves.

Strategies to Engage

- Have students create a Venn diagram highlighting the dual nature of light. Instruct students to label the circle on the left "Wave", the circle on the right "Particle", and the area where the two circles overlap "Both".

Inform students that as they explore the information in this lesson, they will be writing experimental evidence, equations, and explanations in the appropriate places.

Strategies to Explore

- Explain to students that although light travels as a wave, it is actually made up of tiny energy packets, or particles called photons. You can demonstrate this potentially confusing concept by allowing students to observe as you pour salt from one container to another. Point out to students that although it looks like the salt is flowing in a steady stream like a liquid, they know that it is actually made up of separate grains.

Strategies to Extend and Evaluate

- Have students come up with a new term that explains the dual nature of light. Award a prize to the most creative term.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 5.2 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Decide whether each of the following statements is true or false:
   a. Light always behaves like a wave.
   b. Light always behaves like a particle.
   c. Light travels like a particle and gives up its energy like a wave.
   d. Light travels like a wave and gives up its energy like a particle.
2. Which of the following experiments suggested that light was a wave, and which suggested that light was a particle?
   a. the double-slit experiment
   b. the photoelectric effect
   c. black-body radiation
3. Fill in each of the following blanks.
   a. The brightness of a beam of light is determined by the ____________ of the light wave.
   b. The color of a beam of light is determined by the ______________ of the light wave (frequency is also an acceptable answer)
4. What is the name of the quantity depicted by each of the arrows in the diagram below (**Figure** 5.1)? (Source: Sharon Bewick. CC-BY-SA)

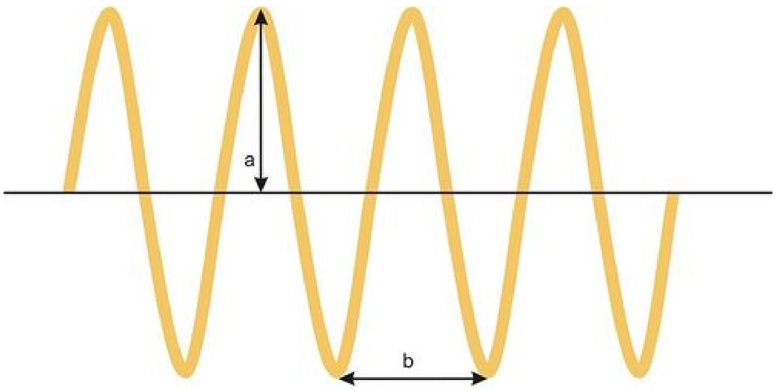


FIGURE 5.1

5. Consider light with a frequency of 4.4×10^{14} Hz. What is the wavelength of this light?
6. What is the frequency of light with a wavelength of 3.4×10^{-9} m?
7. What is the frequency of light with a wavelength of 575 nm?
8. What is the energy of a photon in a beam of light with a frequency of 5.66×10^{8} Hz?

5.3 Light and the Atomic Spectra

Key Concepts

In this lesson students explore continuous and discontinuous spectra.

Lesson Objectives

- Distinguish between continuous and discontinuous spectra.
- Recognize that white light is actually a continuous spectrum of all possible wavelengths of light.
- Recognize that all elements have unique atomic spectra.

Lesson Vocabulary

continuous electromagnetic spectrum A spectrum that contains every possible wavelength of light between the wavelength at the beginning of the list and the wavelength at the end. In the visible range of light, it is a spectrum which contains every possible color between the color at the beginning of the list and the color at the end.

discontinuous electromagnetic spectrum A spectrum that includes some, but not all of the wavelengths in the specified range. In the visible spectrum there are gaps or missing colors.

pure white light A continuous spectrum of all possible wavelengths of light.

atomic spectrum (emission spectrum) A unique, discontinuous spectrum emitted by an element when an electric current is passed through a sample of that element.

Strategies to Engage

- Obtain a wintergreen mint and a pair of pliers. Place the mint between the jaws of the pliers. Turn out the lights and use the pliers to break the mint. Instruct students to note the color of the light emitted. Inform students that in this lesson, they will learn what causes this flash of light. Note: Be sure to practice this demonstration ahead of time because not all wintergreen mints emit light when broken. "When the mint is crushed, electrons in the wintergreen flavor molecules absorb energy, then release it in the form of light."

Strategies to Explore

- Perform the ***Atomic Spectra Viewed Through a Diffraction Grating*** *demonstration. This demonstration is located in the Supplemental Lab Book.*
- Have students complete the lab ***Light and the Atomic Spectra***. This lab is located in the Supplemental Lab Book.

Strategies to Extend and Evaluate

- Have students research the chemistry of fireworks and how fireworks displays relate to atomic emission. Students should be prepared to present their findings to the class.

Lesson Worksheets

There are no worksheets for this lesson.

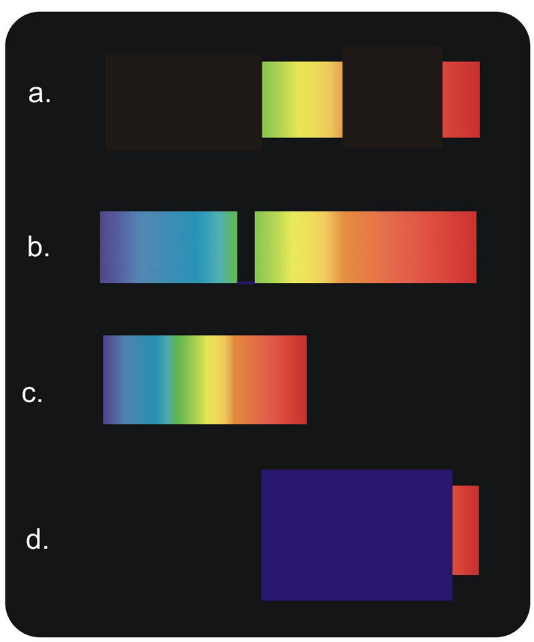
Review Questions

Have students answer the Lesson 5.3 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Decide whether each of the following spectra is continuous or discontinuous. (Source: Sharon Bewick. CC-BY-SA)



2. Decide whether each of the following statements is true or false.
    a. White light has a wavelength of 760 nm.
    b. Sodium's atomic spectrum is an example of a discontinuous spectrum.
    c. Hydrogen's atomic spectrum is an example of a continuous spectrum.
3. Fill in each of the following blanks with the words 'must', 'may' or 'does not.'
    a. A continuous spectrum between 300 nm and 565 nm _____ contain light with a wavelength of 356 nm.
    b. A continuous spectrum between 1000 cm and 1.500 m _____ contain light with a wavelength of 1.234 m.
    c. A discontinuous spectrum between 234 mm and 545 mm ____ contain light with a wavelength of 300 mm.
4. Choose the correct word in each of the following statements.
    a. A continuous spectrum between 532 nm and 894 nm contains light of every wavelength (greater than/less than) 532 nm and (greater than/ less than) 894 nm
    b. A discontinuous spectrum between 532 nm and 894 nm contains light of every wavelength between 532 nm and 894 nm (including/except for) light with a wavelength of 650 nm
5. What is another name for an atomic spectrum?
6. When an electric current is passed through neon, it glows __________.
7. When an electric spark is passed through argon, it glows _________.
8. When an electric current is passed through helium, it glows _________.
9. Suppose you had a solution which contained both dissolved hydrogen and dissolved sodium. If an electric current was passed through this solution, how many lines would you see in the emission spectrum?
10. LEDs, or light emitting diodes, produce light by passing an electric current through a mixture of different atoms (or molecules) and then using their combined emission spectra to light up a room, or a string of Christmas tree lights. Why are white LEDs difficult and expensive to make?

5.4 The Bohr Model

Key Concepts

In this lesson students explore the explanations provided by the Bohr atomic model as well as its limitations.

Lesson Objectives

- Define an energy level in terms of the Bohr model.
- Find the energy of a given Bohr orbit using the equation $E_n = \frac{-R\times h\times c}{n^2}$
- Discuss how the Bohr model can be used to explain atomic spectra.

Lesson Vocabulary

Bohr energy level Distinct energies corresponding to the orbits (or circular paths) of electrons around the atomic nucleus, according to Bohr's model of the atom.

Bohr model of the atom Bohr's explanation of why elements produced discontinuous atomic spectra when struck by an electric current. According to this model, electrons were restricted to specific orbits around the nucleus of the atom in a solar system like manner.

classical physics The laws of physics that describe the interactions of large objects.

quantum mechanics The laws of physics that describe the interactions of very small (atomic or subatomic) objects. Also known as "wave mechanics" and "quantum physics".

Strategies to Engage

- Facilitate a discussion of the changes in the model of the atom over time. Focus the discussion on the location of the electron in the different atomic models. Explain to students that in this lesson they will explore electron arrangements in atoms.

Strategies to Explore

- Emphasize for students that, although Bohr's concept of electrons moving in fixed orbits has since proven to be incorrect, Bohr did discover two very important concepts that are known to be true: electrons occupy specific energy levels within an atom, and energy is quantized.

Strategies to Extend and Evaluate

- Outline the main concepts of the lesson as a class. Discuss the main concepts as you prepare the outline.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 5.4 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Decide whether each of the following statements is true or false:
   a. Niels Bohr suggested that the electrons in an atom were restricted to specific orbits and thus could only have certain energies.
   b. Bohr's model of the atom can be used to accurately predict the emission spectrum of hydrogen.
   c. Bohr's model of the atom can be used to accurately predict the emission spectrum of neon.
   d. According to the Bohr model, electrons have more or less energy depending on how far around an orbit they have traveled.
2. According to the Bohr model, electrons in an atom can only have certain, allowable energies. As a result, we say that the energies of these electrons are _______.
3. The Bohr model accurately predicts the emission spectra of atoms with...
   a. less than 1 electron.
   b. less than 2 electrons.
   c. less than 3 electrons.
   d. less than 4 electrons.
4. Consider an He^+ atom. Like the hydrogen atom, the He^+ atom only contains 1 electron, and thus can be described by the Bohr model. Fill in the blanks in the following statements.
   a. An electron falling from the $n = 2$ orbit of He^+ to the $n = 1$ orbit of He^+ releases ______ energy than an electron falling from the $n = 3$ orbit of He^+ to the $n = 1$ orbit of He^+.
   b. An electron falling from the $n = 2$ orbit of He^+ to the $n = 1$ orbit of He^+ produces light with a ______ wavelength than the light produced by an electron falling from the $n = 3$ orbit of He^+ to the $n = 1$ orbit of He^+.
   c. An electron falling from the $n = 2$ orbit of He^+ to the $n = 1$ orbit of He^+ produces light with a ______ frequency than the light produced by an electron falling from the $n = 3$ orbit of He^+ to the $n = 1$ orbit of He^+.
5. According to the Bohr model, higher energy orbits are located (closer to/further from) the atomic nucleus. This makes sense since negative electrons are (attracted to/repelled from) the positive protons in the nucleus, meaning it must take energy to move the electrons (away from/towards) the nucleus of the atom.
6. According to the Bohr model, what is the energy of an electron in the first Bohr orbit of hydrogen?
7. According to the Bohr model, what is the energy of an electron in the tenth Bohr orbit of hydrogen?

8. According to the Bohr model, what is the energy of an electron in the seventh Bohr orbit of hydrogen?
9. If an electron in a hydrogen atom has an energy of -6.06×10^{-20} J, which Bohr orbit is it in?
10. If an electron in a hydrogen atom has an energy of -2.69×10^{-20} J, which Bohr orbit is it in?
11. If an electron falls from the 5^{th} Bohr orbital of hydrogen to the 3^{rd} Bohr orbital of hydrogen, how much energy is released (you can give the energy as a positive number)?
12. If an electron falls from the 6^{th} Bohr orbital of hydrogen to the 3^{rd} Bohr orbital of hydrogen, what wavelength of light is emitted? Is this in the visible light range?

5.5 References

1. Sharon Bewick. . CC-BY-SA

CHAPTER 6 Quantum Mechanics Model of the Atom TE

Chapter Outline

Quantum Mechanical Model of the Atom

Outline

The chapter *Quantum Mechanical Model of the Atom* consists of five lessons that cover the quantum mechanical model of the atom, energy waves, standing waves, Heisenberg's uncertainty principle, and Schrodinger's equation. Quantum numbers, energy levels, energy sub-levels, and orbital shapes are introduced.

- **Lesson 6.1 The Wave-Particle Duality**
- **Lesson 6.2 Schrodinger's Wave Functions**
- **Lesson 6.3 Heisenberg's Contribution**
- **Lesson 6.4 Quantum Numbers**
- **Lesson 6.5 Shapes of Atomic Orbitals**

Overview

In these lessons, students will explore:

- The wave and particle properties of electrons.
- Electron wave functions and electron density.
- The Heisenberg uncertainty principle.
- The quantum numbers n, ℓ, and m.
- Atomic orbitals.

Science Background Information

This information is provided for teachers who are just beginning to instruct in this subject area.

- **Spin State and Nuclear Magnetic Resonance Imaging**

While the concepts in the study of quantum mechanics may seem elusive to some students, many may be more familiar with one application based on the spin states of the hydrogen atom: Magnetic Resonance Imaging. MRI technology is based on the differing spin states of the hydrogen atom, usually those associated with biological water and fat molecules, and their interaction with strong magnetic fields and radiofrequency waves.

Atoms with an odd number of protons or neutrons in their nucleus possess an intrinsic spin that is quantized. There are two magnetic spin states for the hydrogen nucleus, which can be described as the opposite types of physical spinning: clockwise ($+\frac{1}{2}$) and counterclockwise spinning ($-\frac{1}{2}$). Under normal circumstances, a collection of hydrogen nuclei would display random alignment and both spin states would be equal in energy (degenerate). In the presence of an external magnetic field, however, the hydrogen nuclei can orient with or against the magnetic field, with more nuclei lining up with the magnetic field at a lower energy value. Those nuclei with spins opposing the magnetic field would then be higher in energy by a value of E. If that precise amount of energy is added to the system, in the form of radio waves applied at right angles to the magnetic field, it can cause the lower energy nuclei to absorb and perform a "spin flip" to the higher energy configuration. The radio frequency must match or be in resonance with the nucleus' natural spin. As the "flipped" nuclei gradually relax and realign with the magnetic field, they resume their lower energy spin states and release the absorbed energy. The relaxation rates are a function of the interaction of the nucleus and its physical environment. Incorporation of adjustable magnetic fields can generate a map of very slight difference in resonance frequencies, and thus produce the magnetic resonance image, allowing practitioners to construct images of body tissues with a three-dimensional quality.

The introduction of Magnetic Resonance Imaging has provided a diagnostic revolution in the medical world. Unlike X-rays and CT (computer tomography) scans, no radiation is utilized to produce the image. The key limitations to MRI include eliminating any interference of the magnetic field with metals, thus MRI patients may not have pacemakers, insulin pumps or prosthetic implants. Enhancement of the distinction between normal and diseased tissue is often needed as well, and provided by the introduction of contrast agents. These are usually molecules containing paramagnetic ions, such as Gd^{2+}, which has seven unpaired electrons. These agents are administered intravenously and they serve to highlight visualization of tissue by shortening the relaxation time of the nuclei. Other paramagnetic agents, such as iron oxide and manganese agents are also used for certain applications.

MRI examinations are performed by placing the patient inside the bore of a very large magnet. Other obstacles to this diagnostic tool is that some patients experience claustrophobic anxiety inside the magnet, while others have found the time needed and noise incurred during the data acquisition to be uncomfortable. Despite these hindrances, Magnetic Resonance Imaging has emerged to become an impressive tool for the practice of modern medicine.

Pacing the Lessons

Use the table below as a guide for the time required to teach the lessons of *Quantum Mechanical Model of the Atom*.

TABLE 6.1: Class Periods per Lesson

Lesson	Number of 60 Minute Class Periods
6.1 *The Wave-Particle Duality*	1.0
6.2 *Schroedinger's Wave Functions*	1.0
6.3 *Heisenberg's Contribution*	1.0
6.4 *Quantum Numbers*	1.5
6.5 *Shapes of Atomic Orbitals*	1.0

Managing Materials

The following items are needed to teach the strategies and activities described in the Teachers Edition of the FlexBook for *Quantum Mechanical Model of the Atom*.

TABLE 6.2: The Quantum Mechanical Model of the Atom Materials List

Lesson	Strategy or Activity	Materials Needed
6.1		
6.2	Exploration Activity	Marker, paper
6.3		
6.4		
6.5		

Multimedia Resources

You may find these additional internet resources helpful when teaching *Quantum Mechanical Model of the Atom*: An interactive tour of the atom http://ParticleAdventure.org/ Quantum mechanics video http://www.teachersdomain.org/resource/phy03.sci.phys.fund.quantum/

Possible Misconceptions

Identify: Students may think that the electron cloud is "crowded" with electrons." This misconception may arise because probability patterns appear to show lots of electrons.

Clarify: Electron probability patterns represent the probability of finding a single electron at any given time.

Promote Understanding: Explain to students that electron probability patterns show that probability of finding an electron in a given location increases, then decreases as the distance from the nucleus increases. Probability patterns show the probability of finding a single electron, not the location of a large number of electrons.

Discuss: At the end of the lesson ask "If I were to take a snapshot of an atom with a single electron, how many dots would the photograph show?" *It would show only one dot.*

Ask: "What do the crowded areas versus the less crowded areas of an electron probability pattern show?" *The crowded area show high probability of finding an electron while the less crowded areas show a lower probability of finding an electron in those locations.*

Making the FlexBook Flexible

An important advantage of the FlexBook is the ability it gives you, the teacher, to select the chapters and lessons that you think are most important for your own classes. You should also consult the standards correlation table that follows when selecting chapters and lessons to include in the FlexBook for your classes.

TABLE 6.3: Standards Addressed by the Lessons in Quantum Mechanical Model of the Atom

Lesson	[California] Standards	[NSES] Standards	[AAAS] Benchmarks
6.1	1g, 1n		
6.2			
6.3			
6.4			
6.5			

6.1 The Wave-Particle Duality

Key Concepts

In this lesson students explore the wave and particle properties of electrons.

Lesson Objectives

- Explain the wave-particle duality of matter.
- Define the de Broglie relationship, and give a general description of how it was derived.
- Use the de Broglie relationship to calculate the wavelength of an object given the object's mass and velocity.

Lesson Vocabulary

wave-particle duality of matter Matter exhibits both particle-like and wave-like properties.

Strategies to Engage

- Have students draw a Bohr atomic model. Review charges, masses, and locations of protons, neutrons, and electrons. Re-emphasize for students that Bohr discovered two very important concepts that are known to be true: electrons occupy specific energy levels within an atom, and energy is quantized. Bohr's concept of electrons moving in fixed orbits has proven to be incorrect. Explain to students that in this chapter they will be introduced to the modern atomic model called the quantum mechanical model.

Strategies to Explore

- Use the equation $E = hf$ and deBroglie's equation for wavelength, $\lambda = \frac{h}{m \times v}$, to explain the relationship between the energy of a wave and the wave's frequency, and the relationship between the mass of an object and the object's wavelength. Also, as you go through each example problem, use the equations to explain the concepts explored in this chapter. This will reduce English Language Learner's reliance on language skills. DI **English Language Learners**

Strategies to Extend and Evaluate

- Challenge interested students to derive deBroglie's equation for the wavelength of a particle from the equation $E = mc^2$ and $E = hf$. The first student to correctly show the derivation may then demonstrate to the class.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 6.1 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. In the last chapter you learned that light has wave-like properties and particle-like properties. Can you think of anything else that might have both wave-like properties and particle-like properties?
2. Decide whether each of the following statements is true or false.
   a. Einstein was the first scientist to propose matter waves.
   b. You can see baseballs diffract when you throw them.
   c. The de Broglie's wave equation can only be applied to matter traveling at the speed of light.
   d. Most matter waves are very small, and that is why scientists didn't realize matter had wave-like properties until the 1920s.
3. Choose the correct word in each of the following statements.
   a. The (*more/less*) massive an object is, the longer its wavelength is
   b. The (*faster/slower*) an object is traveling, the shorter its wavelength is
   c. A particle with a mass of 1.0 g has a (*longer/shorter*) wavelength than a particle with a mass of 3.0 g if both are traveling at the same speed
   d. A baseball moving at 10 m/s has a (*longer/shorter*) wavelength than a baseball moving at 4 m/s
4. Choose the correct word in each of the following statements
   a. An electron has a (*longer/shorter*) wavelength than a proton if both are traveling at the same speed.
   b. An electron wave has a (*higher/lower*) frequency than a proton wave if both particles are traveling at the same speed.
   c. If you want to increase the wavelength of an electron, you should (*slow the electron down/speed the electron up*).
5. Choose the correct statement from the options below. The factors that influence an object's wavelength are...
   a. Only the speed of the object
   b. Only the speed of light
   c. The speed of light and the mass of the object
   d. Only the mass of the object
   e. The speed of the object and the mass of the object
6. Choose the correct statement from the options below.

a. Light behaves only like a wave, and matter behaves only like a particle
b. Light behaves only like a wave, and matter behaves only like a wave
c. Light behaves only like a particle, and matter behaves only like a wave
d. Light behaves like a wave and like a particle, but matter only behaves like a particle
e. Light behaves only like a wave, but matter behaves like a wave and like a particle
f. Light behaves like a wave and like a particle, and matter behaves like a wave and like a particle as well

7. Fill in each of the following blanks.
   a. de Broglie used the equations ___________ and _________ to derive an equation for the wavelength of a matter wave.
   b. Scientists first saw matter waves by looking for them in _________. This was a good idea, because ________ are small enough to have matter waves that can be observed in a laboratory.
8. What is the wavelength of a 5.0 kg bowling ball that rolls down the lane at 2.0 m/s?
9. If you walk through a door at 1.0 m/s, and you weight 120 lbs (or 54 kg), what is your wavelength? (This is also approximately the width of the door that would cause your body to diffract.)
10. A car has a mass of 1250 kg. If the car's wavelength is 2.41×10^{-38} m, at what speed is the car traveling?
11. A bobsled sliding down the run at 14.8 m/s has a wavelength of 1.79×10^{-37} m. What is the total mass of the bobsled?

6.2 Schrodinger's Wave Functions

Key Concepts

In this lesson students explore electron wave functions and electron density.

Lesson Objectives

- Distinguish between traveling and standing waves.
- Explain why electrons form standing waves, and what this means in terms of their energies.
- Define an electron wave function and electron density and relate these terms to the probability of finding an electron at any point in space.

Lesson Vocabulary

traveling waves Waves that travel, or move.

standing waves Waves that do *not* travel, or move. They are formed when two traveling waves, moving in opposite directions at the same speed run into each other and combine.

electron wave function A mathematical expression to describe the magnitude, or 'height' of an electron standing wave at every point in space.

electron density The square of the wave function for the electron, it is related to the probability of finding an electron at a particular point in space.

Strategies to Engage

- Preview the lesson vocabulary and lesson objectives. Have students write ten statements from the objective and vocabulary. For example; Traveling waves are waves that do not move. At the end of the lesson have them evaluate their earlier statements and change their responses when necessary.

Strategies to Explore

- Have students write down the lesson objectives, leaving about 5 or 6 lines of space in between. As you explore the lesson, have students write the "answer" to each objective. DI **Less Proficient Readers**

- Students can use a marker and a target to demonstrate the probability distribution of an electron in relation to the nucleus. Instruct students to place the target on the floor and drop the marker from a specific height directly over the target 50 times. Explain to students that this activity illustrates the probability pattern for a single electron atom as shown in Figure 6.8. Point out to students that the probability of finding the electron increases then decreases as the distance from the nucleus increases.

Strategies to Extend and Evaluate

- Have students evaluate the statements introduced under the engagement section (above) and use this information to make a concept map of the lesson content.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 6.2 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Choose the correct word in each of the following statements.
   a. The (*more/less*) electron density at a given location within the atom the more likely you are to find the electron there.
   b. If there is no electron density at a particular point in space, there is (*no/a high*) chance of finding the electron there.
   c. The higher the probability of finding an electron in a certain spot, the (*more/less*) electron density there will be at that spot.
2. The hydrogen ion, H^+ has no electrons. What is the total amount of electron density in a hydrogen atom?
3. Decide whether each of the following statements is true or false.
   a. Only certain electron standing waves are allowed in any particular atom.
   b. Only certain electron energies are allowed in any particular atom
4. The name for the mathematical expression used to describe an electron standing wave is ________.
5. Choose the correct statement.
   a. Einstein first developed the method of describing electron standing waves with wave functions
   b. Planck first developed the method of describing electron standing waves with wave functions
   c. de Broglie first developed the method of describing electron standing waves with wave functions
   d. Schrödinger first developed the method of describing electron standing waves with wave functions
6. Circle all of the statements below which are correct.
   a. The wave function description of electrons predicts that electrons orbit the nucleus just like planets orbit the sun.
   b. The wave function description of electrons predicts that electron energies are quantized

c. The Bohr model of the atom suggests that electron energies are quantized.

7. Fill in the blanks.
   a. Since only certain values are allowed for the energy of an electron in an atom, we say that electron energies are _________.
   b. Allowed electron energies correspond to ______________ that fit perfectly in the atom.
8. Forbidden electron energies correspond to electron standing waves that ________ in the atom.

6.3 Heisenberg's Contribution

Key Concepts

In this lesson students explore the Heisenberg uncertainty principle.

Lesson Objectives

- Define the Heisenberg Uncertainty Principle.
- Explain what the Heisenberg Uncertainty Principle means in terms of the position and momentum of an electron.
- Explain why the Heisenberg Uncertainty Principle helps to justify the fact that a wave function can only predict the probable location of an electron, and not its exact location.

Lesson Vocabulary

momentum (p) The quantity you get when you multiply an object's mass by it's velocity (which as far as you're concerned is the same as its speed).

Heisenberg's Uncertainty Principle Specific pairs of properties, such as momentum and position, are impossible to measure simultaneously without introducing some uncertainty.

Strategies to Engage

- Use this model of an atom to reveal student misconceptions and answer any questions students may have about electrons' arrangements in atoms. The protons and neutrons of this atom make up its nucleus. Electrons surround the nucleus, but they do no circle them like planets around a star, as this model suggests. Review the contributions of Bohr and de Broglie to the modern (quantum mechanical) model of the atom. Explain to students that in this lesson they will learn about the important contribution of Werner Heisenberg, a student of Niels Bohr.

Strategies to Explore

- Challenge students to re-state Heisenberg's Uncertainty Principle using as few words as possible. Consider awarding a prize to the student with the most concise, yet accurate definition. For example: Measuring the position of a particle disturbs its momentum, and vice versa.

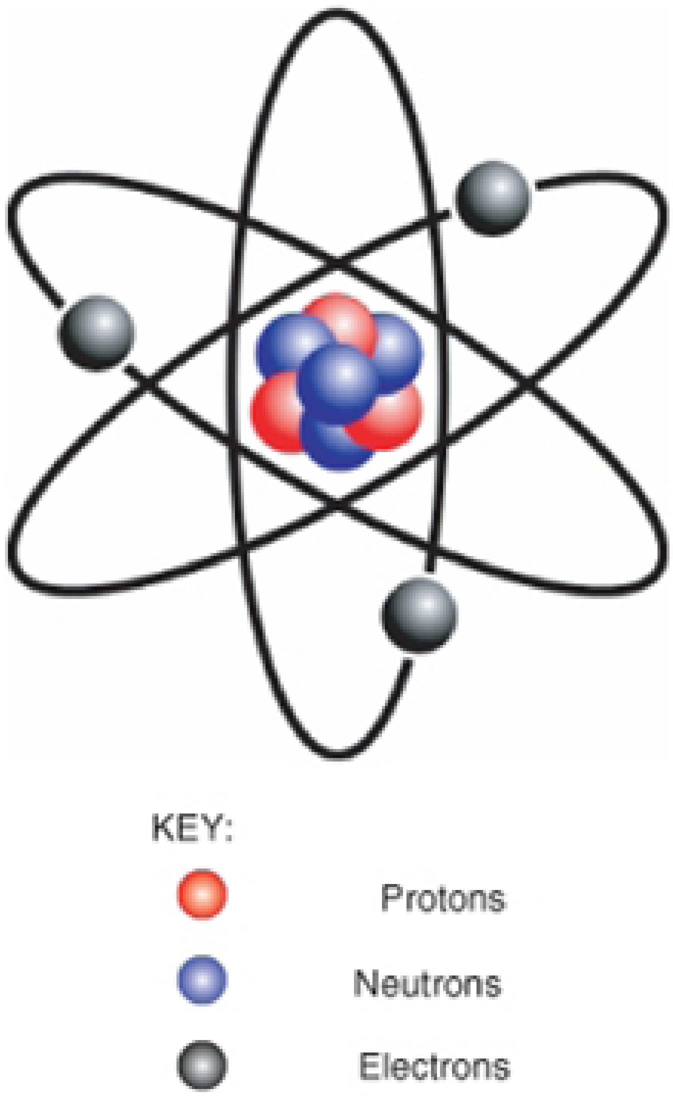


FIGURE 6.1

Strategies to Extend and Evaluate

- Challenge interested students to design a model that would help others understand Heisenberg's Uncertainty Principle. Instruct students to use common objects such as balls and other toys, paper clips, and rulers. Encourage students to share their models with their classmates.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 6.3 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. What types of things in everyday life are impossible to predict with absolute certainty?
2. Why is it impossible to predict the future with absolute certainty?
3. Fill in the blank. According to the _________ Uncertainty Principle, it is impossible to know both an electron's _________ and momentum at the same time.
4. Decide whether each of the following statements is true or false:
    a. According to the Heisenberg Uncertainty Principle, we will eventually be able to measure both an electron's exact position and its exact location at the same time.
    b. The problem that we have when we try to measure an electron's position and its location at the same time is that our measuring equipment is not as good as it could be.
    c. According to the Heisenberg Uncertainty Principle, we cannot know both the exact position and the exact location of a car at the same time.
5. Circle the correct statement The Heisenberg Uncertainty Principle. . .
    a. applies only to very small objects like protons and electrons
    b. applies only to very big objects like cars and airplanes
    c. applies to both very small objects like protons and electrons and very big objects like cars and airplanes

6.4 Quantum Numbers

Key Concepts

In this lesson students explore the quantum numbers n, ℓ, and m.

Lesson Objectives

- Explain the meaning of the principal quantum number, n.
- Explain the meaning of the azimuthal quantum number, ℓ.
- Explain the meaning of the magnetic quantum number, m.

Lesson Vocabulary

quantum numbers Integer numbers assigned to certain quantities in the electron wave function. Because electron standing waves must be continuous and must not 'double over' on themselves, quantum numbers are restricted to integer values.

principal quantum number Defines the energy level of the wave function for an electron, the size of the electron's standing wave, and the number of nodes in that wave.

node A place where the electron wave has zero height. In other words, it is a place where there is no electron density.

azimuthal quantum number Defines the electron sublevel, and determines the shape of the electron wave.

magnetic quantum number Determines the orientation of the electron standing wave in space.

Strategies to Engage

- Point out to students that the probability distribution for an electron in a hydrogen atom has been explored so far. Explain to students that in this lesson they will explore quantum numbers, which are important when it comes to determining the shape of a probability pattern.

Strategies to Explore

- Have students create a chart of the three quantum numbers. As each quantum number is explored and explained, have student fill in the chart with the following information: name, symbol, definition, allowed values, and other important information. Have students save their chart for reviewing lesson content.

Strategies to Extend and Evaluate

- Have students use the information in the chart introduced in the exploration section (above) to create a five-paragraph essay explaining the quantum numbers.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 6.4 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Match each quantum number with the property that they describe.
   a. n - i. shape
   b. ℓ - ii. orientation in space
   c. m_l - iii. number of nodes
2. A point in an electron wave where there is zero electron density is called a _________.
3. Choose the correct word in each of the following statements.
   a. The (*higher/lower*) the value of n, the more nodes there are in the electron standing wave.
   b. The (*higher/lower*) the value of n, the less energy the electron has.
   c. The (*more/less*) energy the electron has, the more nodes there are in its electron standing wave.
4. Fill in the blank. For lower values of n, the electron density is typically found ________ the nucleus of the atom, while for higher values of n, the electron density is typically found __________the nucleus of the atom.
5. Circle all of the statements that make sense: Schrodinger discovered that certain quantities in the electron wave equation had to be integers, because when they weren't, the wave equation described waves which...
   a. were discontinuous
   b. were too small
   c. were too long and narrow
   d. were too short and fat
   e. 'doubled back' on themselves
6. What are the allowed values of ℓ for an electron standing wave with $n = 4$?
7. How many values of ℓ are possible for an electron standing wave with $n = 9$?

8. What are the allowed values of m_l for an electron standing wave with $\ell = 3$?
9. How many different orientations are possible for an electron standing wave with $\ell = 4$?
10. What are the allowed values of m_l for $n = 2$?

6.5 Shapes of Atomic Orbitals

Key Concepts

In this lesson students explore atomic orbitals.

Lesson Objectives

- Define an electron orbital.
- Be able to recognize *s* orbitals by their shape.
- Be able to recognize *p* orbitals by their shape.

Lesson Vocabulary

orbital A wave function for an electron defined by all three quantum numbers, n, ℓ, and m_1. Orbitals define regions in space where there is a high probability of finding the electron.

Strategies to Engage

- Use this opportunity to review the contributions of Bohr, de Broglie, Schrodinger, Born, and Heisenberg to the modern (quantum mechanical) atomic model. Explain to students that in this lesson they will explore atomic orbitals, or the regions in space where there is a high probability of finding the electron.

Strategies to Explore

- Have less proficient readers make a main ideas/details chart as they read the lesson. Instruct them to divide a sheet of paper down the middle and record the main ideas on the left side, and the details for each main idea on the right side. Have students save their chart for reviewing lesson content. DI **Less Proficient Readers**

Strategies to Extend and Evaluate

- Have students create a museum exhibit of the atomic models, principal scientists, experiments, equations, and concepts explored in this chapter.

- Have each student create a poster illustrating the refining of the atomic theory from Democritus to modern atomic theory.
- Have students create a side-by-side comparison of the Bohr model and the quantum mechanical model of the atom.

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled ***Quantum Numbers and Orbital Shapes Worksheet***. Ask the students to complete the worksheet alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 6.5 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Fill in the blanks. When $\ell = 0$, the electron orbital is _________ and when $\ell = 1$, the electron orbital is _________ shaped.
2. The $n = 1$ s orbital has _____ nodes.
3. The $n = 2$ s orbital has _____ nodes.
4. The $n = 2$ p orbital has _____ nodes.
5. The $n = 1$ p orbital has _____ nodes.
6. There are ____ different p orbitals.
7. What energy level (or value of n) has s, p and d orbitals, but no f orbitals?
8. How many different d orbital orientations are there?
9. How many f orbital orientations are there?
10. How many different orbitals are there in the $n = 3$ energy level?

6.6 References

1. . http://commons.wikimedia.org/wiki/Image:Stylised_Lithium_Atom.png. Creative Commons

CHAPTER 7 Electron Configurations for Atoms TE

Chapter Outline

Electron Configurations of Atoms

Outline

This chapter *Electron Configuration of Atoms*, consists of four lessons that cover electron spin, the Aufbau principle, and several methods for indicating electron configuration.

- **Lesson 7.1 The Electron Spin Quantum Number**
- **Lesson 7.2 Pauli Exclusion Principle**
- **Lesson 7.3 Aufbau Principle**
- **Lesson 7.4 Writing Electron Configuration**

Overview

In these lessons, students will explore:

- The electron spin number, its effect on the number of electrons in an orbital and on the magnetic properties of an atom.
- The Aufbau principle, and will use it to predict the orbital in which an electron will be found.
- Orbital representations and electron configurations.

Science Background Information

This information is provided for teachers who are just beginning to instruct in this subject area.

- Orbital Filling Order Exceptions

In assembling the electron configurations for many-electron atoms, one tool that students find valuable is the diagonal rule. This rule provides a guideline that is readily remembered and easily followed to produce accurate electron

configurations for even complicated $d-$ and $f-$ block atoms. One confusing consequence of the diagonal rule is the order of filling the 4s and 3d subshells.

When these orbitals are filled, they are very close in energy. Though as the electrons begin to occupy the empty orbitals, the 4s level is slightly lower in energy than the 3d, thus it is filled first. On the other hand, when both are occupied with electrons, the 4d orbital becomes higher in energy. Thus, in the case that both of these filled levels are composed of valence electrons, the 4s level loses its valence electrons before the 3d level.

The preferential filling of the 4s orbital can also be explained by means of the electron penetration effect. Due to the spherical shape of the s orbital probability density distribution, the likelihood that an electron is found closer to the nucleus is greater than the multi-lobed 3d orbitals.

The similarity in the energy levels of the 4s and 3d orbitals also leads to another interesting consequence. In the electron configuration of the neutral Chromium atom with 24 electrons, the diagonal rule suggests an electron configuration of $1s^22s^22p^63s^23p^64s^23d^4$. The actual electron configuration is $1s^22s^22p^63s^23p^64s^13d^5$, where due to the similarity in energy between the 4s and 3d orbitals, one electron transfers from the 4s to the 3d orbital. The net effect of this exchange yields half-filled 4s and 3d orbitals, and therefore can be justified in terms of generating additional stability. This is also the case for neutral copper atoms, with 29 electrons and a putative electron configuration of $1s^22s^22p^63s^23p^64s^23d^9$. Again as in the example of chromium, an electron transfer occurs, shifting one electron from the 4s orbital to the 3d orbital. For copper, the 4s orbital is now half-filled but added stability is attained by completing the 3d subshell.

The stability afforded to half-filled orbitals is also noted among the $f-$block elements. For example, the electron configuration for Europium (atomic number 63) is $1s^22s^22p^63s^23p^64s^23\text{d}^{10}4p^65s^24d^{10}5p^66s^24f^7$ whereas the next atom, Gadolinium, with atomic number 64, has the additional electron added to the 5d orbital in order to maintain the half-filled stability of the $4f^7$ configuration. The electron configuration for Gadolinium is therefore $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^24f^75d^1$.

The unusual stability of half-filled orbitals can be explained in terms of the disruption afforded by the addition of another electron to this configuration. After the orbital is half-filled, the next additional electron must pair up with another electron, increasing the spin-spin interaction energy and destabilizing the configuration.

Pacing the Lessons

Use the table below as a guide for the time required to teach the lessons of *Electron Configuration of Atoms*.

TABLE 7.1: Class Periods per Lesson

Lesson	Number of 60 Minute Class Periods
7.1 *The Electron Spin Quantum Number*	0.5
7.2 *Pauli Exclusion Principle*	1.0
7.3 *Aufbau Principle*	1.0
7.4 *Writing Electron Configurations*	2.0

Managing Materials

The following items are needed to teach the strategies and activities described in the Teachers Edition of the FlexBook for *Electron Configuration of Atoms*.

TABLE 7.2: Electron Configuration of Atoms Materials List

Lesson	Strategy or Activity	Materials Needed
7.1		
7.2		
7.3		
7.4		

Multimedia Resources

You may find these additional internet resources helpful when teaching *Electron Configuration of Atoms*: Lesson on atomic structure http://www.chemtutor.com/struct.htm Apartment analogy of electron configuration http://kaffee.netfirms.com/Science/activities/Chem/Activity.Electron.Configuration.html

Possible Misconceptions

Making the FlexBook Flexible

An important advantage of the FlexBook is the ability it gives you, the teacher, to select the chapters and lessons that you think are most important for your own classes. You should also consult the standards correlation table that follows when selecting chapters and lessons to include in the FlexBook for your classes.

TABLE 7.3: Standards Addressed by the Lessons in Electron Configuration of Atoms

Lesson	[California] Standards	[NSES] Standards	[AAAS] Benchmarks
7.1	1j		
7.2			
7.3			
7.4	1g		

7.1 The Electron Spin Quantum Number

Key Concepts

In this lesson students explore the electron spin number, its effect on the number of electrons in an orbital and on the magnetic properties of an atom.

Lesson Objectives

- Explain what is meant by the spin quantum number, m_s.
- Explain how the spin quantum number affects the number of electrons in an orbital.
- Explain the difference between diamagnetic atoms and paramagnetic atoms.

Lesson Vocabulary

spin quantum number, The fourth quantum number that must be included in the wave function of an electron in an atom in order to completely describe the electron.

spin-up The term applied to electrons with spin quantum number $m_s = +\frac{1}{2}$.

spin down The term applied to electrons with spin quantum number $m_s = -\frac{1}{2}$.

diamagnetic electrons Two electrons with opposite spins, paired together in an orbital.

diamagnetic atom An atom with no net spin; an atom with *only* diamagnetic electrons.

paramagnetic electron An electron alone in an orbital.

paramagnetic atom An atom with a net spin; an atom with *at least one* paramagnetic electron.

Strategies to Engage

- Take the time to review quantum numbers and how they describe the probable location of the electron, the shape of atomic orbitals, and the orientation of the orbital. Inform students that in this lesson they will explore a fourth quantum number, the spin quantum number, which describes the spin of the electron.

Strategies to Explore

- Have students calculate all possible quantum states for elements 1-10.

Strategies to Extend and Evaluate

- Challenge interested students to come up with a way to explain, model, or illustrate the four quantum numbers. Examples include an analogy to a home address, an explanation that uses a colored chart, or a model that uses balloons. Have students present their information to the class. Facilitate a discussion with students about the limits to the analogies and models.
- Have students complete the lab ***Diamagnetic Levitation***. This lab is located in the Supplemental Lab Book.

Lesson Worksheets

Copy and distribute the worksheet in the *CK-12 Chemistry Workbook* titled ***Quantum Numbers***. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 7.1 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. The principal quantum number describes the size of an electron energy level, the azimuthal quantum number describes the shape of an electron energy level, and the magnetic quantum number describes the orientation of the electron energy level. If there was another quantum number, what do you think it might describe about the electron?
2. There is, in fact, a fourth quantum number that we'll learn about in this lesson. The fourth quantum number is called the spin quantum number. Now can you guess what the final quantum number might describe?
3. Choose the correct statement.
   a. The spin quantum number for an electron can only have the values $m_s = +1$ and $m_s = -1$
   b. The spin quantum number for an electron can only have the value $m_s = 0$
   c. The spin quantum number for an electron can have any integer value between $-\ell$ and $+\ell$
   d. The spin quantum number for an electron can only have the values $m_s = +1/2$ and $m_s = -1/2$
   e. The spin quantum number does not apply to electrons
4. Choose the correct statement.
   a. When two electrons share an orbital, they always have the same spin quantum numbers
   b. When two electrons share an orbital, they always have opposite spin quantum numbers
   c. Two electrons cannot share the same orbital
   d. When two electrons share an orbital there is no way to predict whether or not they will have the same spin quantum numbers
5. Fill in the blanks in the following statement using numbers.

When scientists used the Schrödinger equation with only __ quantum numbers, they found that the Schrödinger equation was pretty good at predicting atomic spectra, except that there were occasionally __ closely spaced lines of light where the Schrödinger equation predicted only __. This led scientists to suggest that a complete description of an electron, which required ___ quantum numbers.

6. In many atomic spectra, there are two very closely spaced lines of light which can only be predicted by including the spin quantum number into the Schrödinger equation. Decide whether the following statements about these two lines are true or false.
    a. the two lines spread further apart when the atom is placed in a magnetic field
    b. the two lines move closer together when the atom is placed in a magnetic field
    c. the two lines are the result of an experimental error. If scientists are careful, they find that there is really just one line.
    d. the two lines actually result from the fact that there are two very closely spaced energy states
7. Goudsmit and Uhlenbeck proposed the existence of
    a. the principal quantum number
    b. the azimuthal quantum number
    c. the spin quantum number
    d. the magnetic quantum number
8. Circle all of the quantum numbers that tell you about the region in space where you're most likely to find the electron.
    a. the spin quantum number
    b. the magnetic quantum number
    c. the principal quantum number
    d. the azimuthal quantum number
9. Select the correct statement from the list below. An electron with a spin quantum number of $m_s = -1/2$
    a. cannot share an orbital with an electron that has a spin quantum number of $m_s = +1/2$
    b. prefers to share an orbital with an electron that has a spin quantum number of $m_s = -1/2$
    c. cannot share an orbital with an electron that has a spin quantum number of $m_s = -1/2$
    d. cannot share an orbital with another electron
10. What is the total spin in an electron orbital if
    a. the orbital contains one 'spin-up' electron
    b. the orbital contains one 'spin-down' electron
    c. the orbital contains two 'spin-up' the orbital contains one 'spin-up' electron and one 'spin-down' electron
    d. .

7.2 Pauli Exclusion

Key Concepts

In this lesson students explore the Pauli exclusion principle and its implication for electron arrangements in atoms.

Lesson Objectives

- Explain the meaning of the Pauli Exclusion Principle.
- Determine whether or not two electrons can coexist in the same atom based on their quantum numbers.
- State the maximum number of electrons that can be found in any orbital.

Lesson Vocabulary

Pauli Exclusion Principle No two electrons may occupy the same quantum state in an atom simultaneously; no two electrons in an atom can have the same four quantum numbers.

Strategies to Engage

Strategies to Explore

- Wolfgang Pauli was a close friend of both Werner Heisenberg and Niels Bohr. Facilitate a discussion with students about some of the conversations these three friends might have had.

Strategies to Extend and Evaluate

Write each of the following statements on the board or on chart paper:

- No two electrons in an atom can have the same four quantum numbers.
- No atomic orbital can contain more than two electrons.
- Electrons in the same atom with the same spin must be in different orbitals.
- Electrons in the same orbital of the same atom must have different spins.

Have students come up with example problems that illustrate each statement.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 7.2 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Electrons in the same orbital must have different spin quantum numbers. What is true of the other three quantum numbers for two electrons in the same orbital?
2. Electrons in different orbitals can have the same spin quantum numbers. What is true of the other three quantum numbers for two electrons in different orbitals?
3. Fill in the blank using either the word 'can', or 'cannot'.
   a. An electron with the quantum numbers $n = 1$, $\ell = 0$, $m_l = 0$ and $m_s = +1/2$ _____ exist in the same atom as an electron with the quantum numbers $n = 2$, $\ell = 0$, $m_l = 0$ and $m_s = +1/2$
   b. An electron with the quantum numbers $n = 1$, $\ell = 0$, $m_l = 0$ and $m_s = +1/2$ _____ exist in the same atom as an electron with the quantum numbers $n = 1$, $\ell = 0$, $m_l = 0$ and $m_s = -1/2$
4. Fill in the blanks using numbers.
   a. There is only 1 orbital at the $n = 1$ energy level. Therefore the $n = 1$ energy level can hold a maximum of __ electrons
   b. There are 4 orbitals at the $n = 2$ energy level. Therefore the $n = 2$ energy level can hold a maximum of __ electrons
   c. There are 9 orbitals at the $n = 3$ energy level. Therefore the $n = 3$ energy level can hold a maximum of __ electrons
   d. There are 16 orbitals at the $n = 4$ energy level. Therefore the $n = 4$ energy level can hold a maximum of __ electrons
5. What is the maximum number of electrons that can exist in p orbitals at energy levels with $n < 3$.
6. What is the maximum number of electrons that can exist in p orbitals at energy levels with $n < 5$.

7.3 Aufbau Principle

Key Concepts

In this lesson students explore the Aufbau Principle and will use it to predict the orbital in which an electron will be found.

Lesson Objectives

- Explain the Aufbau Principle.
- Given two different orbitals, predict which the electron will choose to go into.

Lesson Vocabulary

Aufbau principle Electrons will fill available orbitals starting with those at the lowest energy before moving to those at higher energies.

Strategies to Engage

Strategies to Explore

- Have students come up with more descriptive names for both the Pauli exclusion principle, and the Aufbau Principle.

Strategies to Extend and Evaluate

- Challenge interested students to come up with a skit illustrating the Aufbau Principle. Have them perform the skit for their classmates. Then facilitate a class discussion of the limitations of the skit to accurately represent the Aufbau Principle.
- Use student descriptions to come up with a composite diagram of the arrangement of electrons in energy levels, sublevels, and orbitals. This will give you the opportunity to review lesson concepts and clear up any misconceptions students may have about the lesson content.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 7.3 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. While we have talked about emission spectra, another type of spectra is known as absorption spectra. In emission spectra, the atom emits lines of light like those you saw in the examples of atomic spectra. In absorption spectra, the atom absorbs lines of light, rather than emitting them. Can you explain this in terms of electrons and orbitals? What do you think the relationship between absorption spectra and emission spectra might be?
2. If an electron has a "choice" between going into an orbital in the $n = 1$ energy level or an orbital in the $n = 2$ energy level, which do you think it chooses?
3. If an electron in the $n = 3$ energy level has a "choice" between going into an orbital with $\ell = 0$ or an orbital with $\ell = 1$, which do you think it chooses?
4. Select the correct statement. According to the Aufbau Principle
   a. orbitals with higher values of n fill up first
   b. orbitals in the same energy level, but with higher values of ℓ fill up first
   c. orbitals with lower values of n fill up first
   d. it is impossible to predict which orbitals will fill up first
5. Decide whether each of the following statements is true or false.
   a. Electrons in different orbitals have different energies.
   b. An electron will enter an orbital of higher energy when a lower energy orbital is already filled.
   c. For some atoms the first energy level can contain more than two electrons.
6. Does the electron in the hydrogen atom absorb or emit energy when it makes a transition between the following energy levels:
   a. $n = 2$ to $n = 4$
   b. $n = 6$ to $n = 5$
   c. $n = 3$ to $n = 6$
7. Fill in the blanks. There is one s orbital, three p orbitals, and five d orbitals in the $n = 3$ energy level of an atom. If a particular atom has a total of 5 electrons in the $n = 3$ energy level, then there are...
   a. ___ electrons in the s orbital
   b. ___ electrons in p orbitals
   c. ___ electrons in d orbitals
8. Fill in the blanks. There is one s orbital, three p orbitals, five d orbitals and $7f$ orbitals in the $n = 4$ energy level of an atom. If a particular atom has a total of 7 electrons in the $n = 4$ energy level, then there are...
   a. ___ electrons in the s orbital
   b. ___ electrons in p orbitals
   c. ___ electrons in d orbitals
   d. ___ electrons in f orbitals

9. According to the Aufbau rule, which of the following atoms has a sub-shell that is exactly half-filled?
    a. *Ba*
    b. *Al*
    c. *C*
    d. *As*
    e. *O*

7.4 Writing Electron Configurations

Key Concepts

In this lesson students will explore orbital representations and electron configurations.

Lesson Objectives

- Figure out how many electrons can exist at any given sublevel.
- Figure out how many different sublevels can exist at any given energy level.
- Be able to write the electron configuration of any element, given the total number of electrons in that element.
- Be able to write either orbital representations or electron configuration codes.

Lesson Vocabulary

electron configuration A short hand notation to indicate the electron orbitals which are filled in a particular atom.

diagonal rule The electrons fill orbitals in order of increasing 'quantum number sum' $(n+\ell)$. When two orbitals share the same 'quantum number sum', they will be filled in order of increasing n.

quantum number sum The sum of the principal quantum number, n, and the azimuthal quantum number, ℓ, for an electron. That is $n+\ell$.

Strategies to Engage

- Write $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$ on the board. Tell students that what you have just written is the electron configuration for Calcium, which shows the way the electrons are arranged in a Calcium atom. Ask students if they can figure out what the letters and numbers represent. Inform students that in this lesson they will not only learn what those letters and numbers represent, but they will also learn how to write the electron configuration for atoms of elements.

Strategies to Explore

- Emphasize for English Language Learners the figures in this lesson and use them to teach important concepts. Have a language proficient student "read" each visual, pointing out important concepts. **English Language Learners**

Strategies to Extend and Evaluate

- Have students play a game of "electron configuration bingo". Instruct students to draw a five by five grid on a sheet of notebook paper and mark a "free space" in the middle. Have them write any of the first 24 elements in any order into the remaining spaces. Read electron configurations and have students use a highlighter or pen to mark off the elements it represents. The first student to correctly mark five elements horizontally, vertically, or diagonally wins.
- Have students research humorous stories about Wolfgang Pauli. Students should be prepared to share their findings with the class.

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Orbital Representations Worksheet**. Ask the students to complete the worksheet alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 7.4 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Write the electron configuration for beryllium. Beryllium has 4 electrons.
2. Write the electron configuration for silicon. Silicon has 4 electrons.
3. Write the electron configuration for nitrogen. Nitrogen has 7 electrons.
4. Write the electron configuration for chromium. Chromium has 24 electrons.
5. Write the electron configuration for silver. Silver has 47 electrons.

CHAPTER 8 Electron Configurations and the Periodic Table TE

Chapter Outline

Unit 3 Periodic Relationships

Outline

This unit, Periodic Relationships, includes three chapters that explore the periodic table.

- **Chapter 8 Electron Configurations and The Periodic Table**
- **Chapter 9 Relationships Between the Elements**
- **Chapter 10 Trends on the Periodic Table**

Overview

Electron Configuration and the Periodic Table

This chapter develops the relationship between the electron configuration of atoms and their positions on the periodic table.

Relationships Between the Elements

This chapter introduces the chemical families caused by electron configuration, the concept of valence electrons, and Lewis electron dot formulas.

Trends on the Periodic Table

This chapter explains the periodic change in atomic size and its relationship to the periodic trends for ionization energy and electron affinity.

Electron Configurations and the Periodic Table

Outline

The chapter *Electron Configurations and the Periodic Table* consists of three lessons that explore the relationship between the electron configuration of an element and its position on the periodic table.

- **Lesson 8.1 Electron Configurations of Main Group Elements**
- **Lesson 8.2 Orbital Configurations**
- **Lesson 8.3 The Periodic Table and Electron Configurations**

Overview

In these lessons, students will explore:

- The relationship between the number of valence electrons an element has, and its position on the periodic table.
- Hund's rule and then use it to write orbital representations for elements.
- The relationship between an element's position on the periodic table and its atom's highest occupied energy level.

Science Background Information

This information is provided for teachers who are just beginning to instruct in this subject area.

- The Upper Limit of the Periodic Table

The Periodic Table has been acknowledged as one of the most influential keys to understanding modern chemistry. A wealth of information is organized into a readily interpretable array of essential atomic data. Since the days of Dmitri Mendeleev, who is credited with arranging our modern periodic table on the basis of physical similarities, atomic physicists have drastically extended the number of elements by the preparation of artificial elements. These are atoms not found naturally on Earth due to radioactive decay instability but have been created synthetically by atomic bombardment and collisions.

The very first synthetic element was the result of many years of searching for the elusive missing element to be inserted between molybdenum and ruthenium, an omission noted and a space left open by Mendeleev. Many efforts claiming to have identified element 43 were made but not substantiated. Conclusive evidence for the production of a new element was made by Emilio Segré and Carlo Perrier in 1937 after they collided molybdenum atoms with the heavy isotope of hydrogen known as deuterium. Later trace amounts of technetium were identified among the decay products of uranium fission. The name technetium was chosen from the Greek word for artificial.

The next synthetic element, 61, promethium, was produced by a similar method. Jakob Marinsky and Larry Glendenin at MIT bombarded neodymium atoms with neutrons obtained as byproducts of uranium decay. Their 1946 announcement named the new element after the mythological Prometheus, who, according to legend was responsible for bringing fire to mankind.

The decade of the 1940's also marked the creation of the first trans-uranium element. Neptunium was the result of Berkeley scientists Edwin McMillan and Philip Abelson colliding uranium with neutrons as was the concurrent production of element 94, named plutonium in the sequence correlating with the modern group of solar system planets. One name suggested for element 94 was "extremium" offering the proposition that this artificially produced element was the upper limit or heaviest possible atom.

Since that time, the quest for producing super-heavy elements has continued with the question of where and when that upper limit, if it exists, will be reached. Currently, (2009) the as-yet unnamed Element 118, a member of the noble gas family, maintains its status as the heaviest element. Three atoms of element 118 were reportedly created by fusing californium atoms with calcium atoms in 2006 at Lawrence Livermore Laboratory. In the last year, claims

suggesting the existence of Element 122 have also been reported but as yet, experimental replications have failed to reproduce this evidence.

Is there an upper limit to the periodic table? The intrinsic instability with respect to nuclear decay appears to limit the production of elements with atomic numbers greater than that of uranium. Most of the trans-uranium elements have extremely short half-lives and very limited production quantities. Attempting to load the tiny atomic nucleus with 100+ protons appears to provide a barrier that may have reached its synthetic limit.

Pacing the Lessons

Use the table below as a guide for the time required to teach the lessons of *Electron Configurations and the Periodic Table*.

TABLE 8.1: Class Periods per Lesson

Lesson	Number of 60 Minute Class Periods
8.1 *Electron Configurations of Main Group Elements*	1.0
8.2 *Orbital Configurations*	1.5
8.3 *The Periodic Table and Electron Configurations*	1.0

Managing Materials

The following items are needed to teach the strategies and activities described in the Teachers Edition of the FlexBook for *Electron Configurations and the Periodic Table*.

TABLE 8.2: Electron Configurations and the Periodic Table Materials List

Lesson	Strategy or Activity	Materials Needed
8.1		
8.2		
8.3		

Multimedia Resources

You may find these additional web based resources helpful when teaching *Electron Configurations and the Periodic Table*:

Interactive periodic table with basic information about each element http://www.chemicalelements.com/

Possible Misconceptions

Making the FlexBook Flexible

An important advantage of the FlexBook is the ability it gives you, the teacher, to select the chapters and lessons that you think are most important for your own classes. You should also consult the standards correlation table that follows when selecting chapters and lessons to include in the FlexBook for your classes.

TABLE 8.3: Standards Addressed by the Lessons in Electron Configurations and the Periodic Table

Lesson	[California] Standards	[NSES] Standards	[AAAS] Benchmarks
8.1	1d		
8.2	1d		
8.3			

8.1 Electron Configurations of Main Group Elements

Key Concepts

In this lesson, students explore the relationship between the number of valence electrons and element has and its position on the periodic table.

Lesson Objectives

- Explain how the elements in the Periodic Table are organized into rows and columns.
- Explain how the electron configurations within a column are similar to each other.

Lesson Vocabulary

Periodic Table Scientists use the Periodic Table to summarize what they know about the existing elements. Elements of similar size are found in the same row, while elements with similar chemical properties are found in the same column.

chemical properties The ways in which an element reacts with another element or compound.

valence electrons The electrons in an atom with the highest value of n (the electrons in the highest energy level).

non-valence electrons All electrons in atom which are not valence electrons. Non-valence electrons are not important in determining an element's chemical properties because they rarely get involved in chemical reactions.

alkali metals Group 1A metals. These are elements found in the first column of the Periodic Table, excluding hydrogen.

alkaline earth metals Group 2A metals. These are elements found in the second column of the Periodic Table.

noble gases Group 8A elements. These are elements found in the eight column of the Periodic Table. They are inert, which means that they are very non-reactive.

Strategies to Engage

- Have students research the Hindenberg disaster of 1939. Point out to students that this disaster could have been avoided had the airship been filled with helium instead of hydrogen. Inform students that in this lesson,

they will learn how electron configurations can be used to predict the properties of elements including their ability to react with each other.

Strategies to Explore

- This lesson provides the opportunity to introduce students to the concept of bonding. For example, when discussing the alkali metals, point out to students that they are very reactive. Ask them to come up with reasons as to why the alkali metals are so reactive. Don't be afraid to informally introduce the octet rule.

Strategies to Extend and Evaluate

- Have students write a matching quiz of the vocabulary explored in this lesson. Instruct students to trade quizzes with another student. Have students grade the quizzes in groups of four.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 8.1 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Take a look at the Periodic Table. How would you describe it? Why do you think it has such a funny shape?
2. Can you suggest how elements in the same column of the Periodic Table might be similar?
3. Choose the correct statement.
   a. Mg has only 1 valence electron in an s orbital
   b. F has only 1 valence electron in an s orbital
   c. O has only 1 valence electron in an s orbital
   d. Kr has only 1 valence electron in an s orbital
   e. Fr has only 1 valence electron in an s orbital
4. Circle the appropriate element for each blank.
   a. _____________ (Mg/N) has 2 valence electrons in an s orbital, and 3 valence electrons in p orbitals.
   b. _____________ (As/B) has 2 valence electrons in an s orbital, and 3 valence electrons in p orbitals
   c. _____________ $(Cl/P/Li)$ has 2 valence electrons in an s orbital, and 5 valence electrons in p orbitals
   d. _____________ $(Al/Li/Na)$ has 1 valence electron in a p orbital
5. Choose the correct statement.
   a. Group 1A elements have a total of 3 valence electrons
   b. Group 5A elements have a total of 2 valence electrons
   c. Group 7A elements have a total of 4 valence electrons

d. Group 8A elements have a total of 8 valence electrons
e. Group 2A elements have a total of 5 valence electrons
f. Group 1A elements have a total of 3 valence electrons

6. Fill in the blanks.

a. *N* has ___ valence electrons in an *s* orbital
b. *N* has ___ valence electrons in *p* orbitals
c. *N* has a total of ___ valence electrons
d. *Ca* has ___ valence electrons in *s* orbitals
e. *Ca* has ___ valence electrons in *p* orbitals
f. *Ca* has a total of ___ valence electrons

7. Decide whether each of the following statements is true or false.

a. *K* has 1 valence electron in an *s* orbital
b. *Ge* has 2 valence electrons in an *s* orbital
c. *Se* has 4 valence electrons in *p* orbitals
d. *B* has 3 valence electrons in *p* orbitals
e. *F* has 2 valence electrons in an *s* orbital, and 7 valence electrons in *p* orbitals
f. *Ca* has a total of 4 valence electrons

8. Match the element to its valence electrons.

a. *Sr* - i. a total of 8 valence electrons
b. *I* - ii. a total of 2 valence electrons
c. *Ne* - iii. a total of 5 valence electrons
d. *N* - iv. a total of 7 valence electrons

9. Fill in the blanks.

a. *Ba* has __ valence electron(s) in an *s* orbital, and __ valence electron(s) in *p* orbitals
b. *Sn* has __ valence electron(s) in an *s* orbital, and __ valence electron(s) in *p* orbitals
c. *S* has __ valence electron(s) in an *s* orbital, and __ valence electron(s) in *p* orbitals
d. *Po* has __ valence electron(s) in an *s* orbital, and __ valence electron(s) in *p* orbitals
e. *Na* has __ valence electron(s) in an *s* orbital, and __ valence electron(s) in *p* orbitals

10. List all of the elements with exactly 2 valence electrons in *p* orbitals.
11. An element has 2 valence electrons in an *s* orbital and 4 valence electrons in *p* orbitals. If the element is in the second row of the Periodic Table, which element is it?
12. An element has 2 valence electrons in an *s* orbital and 6 valence electrons in *p* orbitals. If the element is in the same row as In, which element is it?

8.2 Orbital Configurations

Key Concepts

In this lesson students explore Hund's rule and use it to write orbital representations for elements.

Lesson Objectives

- Draw orbital diagrams.
- Define Hund's Rule.
- Use Hund's Rule to decide how electrons fill sublevels with more than one orbital.

Lesson Vocabulary

orbital diagram Orbital diagrams are drawn by representing each orbital as a box, each 'spin-up' electron in an orbital as an upward pointing arrow in the box, and each 'spin-down' electron in an orbital as a downward pointing arrow in the box.

Hund's rule Every orbital in a sublevel is singly occupied before any orbital is doubly occupied. All of the electrons in singly occupied orbitals have the same spin.

Strategies to Engage

- Begin with a discussion of how electrons behave in atoms. Have each student write down two or three thoughts they may have about electrons and how they behave in atoms. Facilitate a discussion with students and address any misconceptions that may become evident at this time.

Strategies to Explore

Strategies to Extend and Evaluate

- In order to assess student understanding, have them draw orbital diagrams that violate the Aufbau Principle, Pauli Exclusion Principle, and Hund's rule. Students should then explain each rule using their drawings. DI **English Language Learners**

- Challenge interested students to research the relationship between electron arrangement and the ability of some elements to behave as semiconductors. Students should be prepared to share their findings with the rest of the class.

Lesson Worksheets

There are no worksheets for this lesson.

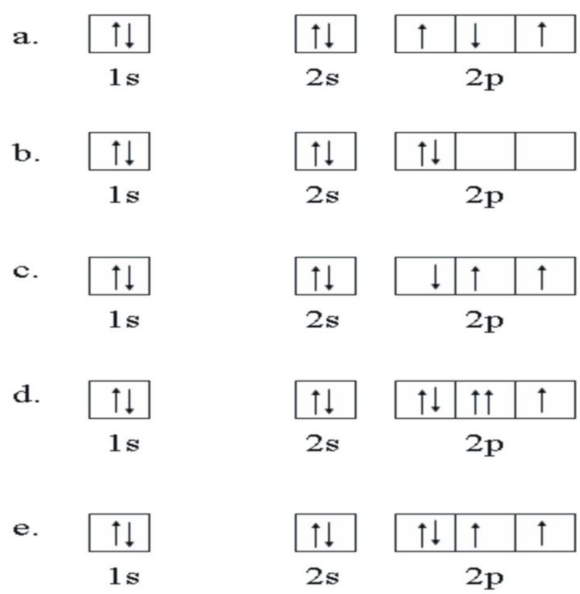
Review Questions

Have students answer the Lesson 8.2 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Which of the following is a valid orbital diagram?



2. Draw the orbital diagram for lithium (Li).
3. Draw the orbital diagram for carbon (C).
4. Draw the orbital diagram for fluorine (F).
5. Draw the orbital diagram for oxygen, O. Use it to answer the following questions:
    a. an oxygen atom has ___ unpaired valence electrons
    b. an oxygen atom has ___ paired valence electrons
    c. an oxygen atom has ___ paired non-valence electrons
    d. an oxygen atom has ___ unpaired non-valence electrons
6. Draw the orbital diagram for neon, Ne. Use it to answer the following questions:
    a. a neon atom has ___ unpaired valence electrons
    b. a neon atom has ___ paired valence electrons
    c. a neon atom has ___ paired non-valence electrons
    d. a neon atom has ___ unpaired non-valence electrons

7. Decide whether each of the following statements is true or false.
   a. Every orbital in a sublevel is doubly occupied before any orbital is singly occupied.
   b. Every orbital in a sublevel is singly occupied before any orbital is doubly occupied.
   c. All electrons in singly occupied orbitals have the same spin.
   d. The two electrons in a single orbital have the same spin.
   e. All electrons in singly occupied orbitals have different spins.
   f. The two electrons in a single orbital have different spins.
8. Draw the orbital diagram for phosphorus P.
9. Draw an orbital diagram for silicon, Si. Use it to answer the following questions:
   a. a silicon atom has ___ unpaired valence electrons
   b. a silicon atom has ___ paired valence electrons
   c. a silicon atom has ___ paired non-valence electrons
   d. a silicon atom has ___ unpaired non-valence electrons
10. Draw an orbital diagram for Mn. Use it to determine the total number of unpaired electrons in an Mn atom.

8.3 The Periodic Table and Electron Configurations

Key Concepts

In this lesson students explore the relationship between an element's position on the periodic table and its atom's highest occupied energy level.

Lesson Objectives

- Relate an element's position in the PT to the energy level of its valence electrons. (Excluding transition metals, lanthanides and actinides.)
- Relate an element's position in the PT to the sublevel of its highest energy valence electrons.
- Explain why there are only two elements in the first row of the PT.

Lesson Vocabulary

transition metals Elements in the *d* sublevel block (columns 1-B through 8B) of the Periodic Table. The highest energy electrons in transition metals are found in d orbitals.

lanthanides and actinides Elements in the *f* sublevel block of the Periodic Table. The highest energy electrons in lanthanides and actinides are found in *f* orbitals.

The elements in the Periodic Table in columns 1A and 2A (excluding hydrogen). All valence electrons for elements in the *s* sublevel block are in *s* orbitals.

The elements in the Periodic Table in columns 3A through 8A (excluding helium). The highest energy valence electrons for elements in the *p* sublevel block are in *p* orbitals.

The elements in the Periodic Table in columns 1B through 8B (also known as transition metals).

The elements in the lanthanide and actinide rows of the Periodic Table.

noble gases Group 8A elements. These are elements found in the eight column of the Periodic Table. They are inert, which means that they are very non-reactive.

inert Non-reactive.

Strategies to Engage

- Have students read the review questions at the end of this section. This way, students will be familiar with the types of information that they will explore in this section.

Strategies to Explore

- Have each student choose a group of main group elements. Instruct students to write the electron configuration for the first four elements in a vertical column and write down two patterns they observe. (They all have the same number of valence electrons. The energy level of the valence electrons increases as you go down the group of elements.) Have students team up with two classmates who have chosen different group of elements and write down patterns they observe among the groups. (The same patterns appear in each group. The number of valence electrons is equal to the group number.)

Strategies to Extend and Evaluate

- Have students research and describe the physical properties of elements in their chosen group from the exploration section (above). Instruct students to write a short paragraph about the chemical similarities of the elements in the group and explain in terms of electron configuration, why the elements have those properties.

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titles **The Periodic Table and Electron Configurations**. Ask the students to complete the worksheet alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 8.3 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Use the Periodic Table to determine the energy level of the valence electrons in each of the following elements.
   a. B
   b. Ga
   c. Rb
   d. At
   e. He
2. Fill in the blanks:
   a. B is in the __ level block of the Periodic Table
   b. Sr is in the __ level block of the Periodic Table

c. Fe is in the __ level block of the Periodic Table
d. Cs is in the __ level block of the Periodic Table
e. O is in the __ level block of the Periodic Table

3. Use the Periodic Table to determine the energy level and sublevel of the highest energy electrons in each of the following elements:
   a. N
   b. Ca
   c. Rb
   d. P
   e. In
4. Decide whether each of the following statements is true or false.
   a. Li has valence electrons in the $n = 1$ energy level.
   b. Si has valence electrons in the $n = 3$ energy level.
   c. Ga has valence electrons in the $n = 3$ energy level.
   d. Xe has valence electrons in the $n = 5$ energy level.
   e. P has valence electrons in the $n = 2$ energy level.
5. Match the element to the sublevel block it is found in:
   a. C - i. s sublevel block
   b. Cs - ii. p sublevel block
   c. Ce - iii. d sublevel block
   d. Cr - iv. f sublevel block
6. The first row of the Periodic Table has:
   a. 1 element
   b. 2 elements
   c. 3 elements
   d. 4 elements
   e. 5 elements
7. Use the Periodic Table to determine which of the following elements has the highest energy valence electrons.
   a. Sr
   b. As
   c. H
   d. At
   e. Na
8. Use the Periodic Table to determine which of the following elements has the lowest energy valence electrons.
   a. Ga
   b. B
   c. Cs
   d. Bi
   e. Cl
9. Which energy level does the first row in the d sublevel block correspond to?

Chapter 9 Relationships Between the Elements TE

Chapter Outline

Relationships Between the Elements

Outline

The chapter Relationships Between the Elements consists of six lessons that introduce the chemical families caused by electron configuration, the concept of valence electrons, and Lewis electron dot formulas.

- **Lesson 9.1 Families on the Periodic Table**
- **Lesson 9.2 Electron Configurations**
- **Lesson 9.3 Lewis Dot Electron Diagrams**
- **Lesson 9.4 Chemical Family Members Have Similar Properties**
- **Lesson 9.5 Transition Elements**
- **Lesson 9.6 Lanthanide and Actinide Series**

Overview

In these lessons, students will explore:

- The electron configurations of families of elements.
- A shortcut method for writing electron configurations.
- Electron dot diagrams.
- Trends in chemical reactivity within chemical families.
- Electron configurations of transition elements.
- The electron configurations of lanthanides and actinides.

Science Background Information

This background information is provided for teachers who are just beginning to teach in this subject area.

- The Discovery of the Noble Gases

The Group 8A elements are known both as the noble gases and with respect to their lack of reactivity, the inert gases. This unifying characteristic of this group of elements can be explained by the modern consideration of their electron configuration, specifically their filled valence shells.

Their lack of participation in chemical reactivity and bonding hindered the progress of identifying these elements. Although the first isolation of what ultimately become as argon was accomplished by Henry Cavendish in 1785, it was not conclusively shown to be a single type of atom until 1894. Cavendish removed the nitrogen, which he knew as "phlogisticated air" and oxygen, but noticed a persistent amount of residual gas. Rayleigh and Ramsay noticed discrepancies in the density of nitrogen gas measured by different mechanism and were the first to isolate a noble gas. For their efforts, each of them were awarded Nobel Prizes

Despite its presence as the second most abundant element in the universe, the existence of the element helium was not suspected until 1868, when in the solar spectrum, emission lines were discovered by French astronomer Pierre Janssen which did not correspond to lines for any previously known elements. The unusual new substance was called helium because it was identified in the solar spectrum before it was found on Earth. William Ramsay was also involved in isolating helium, in this case from uranium salts treated with strong acid. Not long after Ramsay's work, Kansas geologists found an unidentifiable gas in the mixture from an oil-drilling operation, which was later measured to be helium. Currently, most available helium gas is obtained from extraction of natural gas.

The discovery of the element neon, also credited to William Ramsay and Morris Travers, occurred with considerable excitement. This inert gas was obtained after removing nitrogen, oxygen and argon from a sample of liquefied air and his team happened to heat the residual gas sample. The gas unexpectedly yielded a bright red glow, now familiar as a neon sign.

In the same series of experiments that produced neon, krypton and xenon were also identified by Ramsay and Travers in 1898. Krypton's name was borrowed for use in the comic books about Superman to designate the fictional substance that their hero was vulnerable to. Xenon, like the other members of this chemical family, remained characteristically inert until 1962, when chemist Neil Bartlett found that platinum hexafluoride salts react with this previously unreactive gas. Since that time, other compounds containing xenon have been prepared including Xenon tetroxide, XeO_4, and xenon difluoride, XeF_2. A limited number of krypton compounds have also been reported such as KrF_2.

The discovery of the remaining member of this chemical family, radon, was in part, the result of the research of Pierre and Marie Curie. They were responsible for isolating the radioactive elements polonium and radium. A German physicist, Friedrich Dorn, found that when radium is exposed to air, another radioactive gas was produced. This gas was further characterized by William Ramsay, who, in 1903 determined its atomic weight and suggested its placement among the noble gases.

Pacing the Lessons

Use the table below as a guide for the time required to teach the lessons of *Relationships Between the Elements.*

TABLE 9.1: Class Periods per Lesson

Lesson	Number of 60 Minute Class Periods
9.1 *Families on the Periodic Table*	0.5
9.2 *Electron Configurations*	0.5
9.3 *Lewis Electron Dot Diagrams*	0.5
9.4 *Chemical Family Members Have Similar Properties*	0.5
9.5 *Transition Elements*	0.5

TABLE 9.1: (continued)

Lesson	Number of 60 Minute Class Periods
9.6 *Lanthanide and Actinide Series*	0.5

Managing Materials

The following materials are needed to teach the strategies and activities described in the Teachers Edition of the FlexBook for *Relationships Between the Elements.*

TABLE 9.2: Relationships Between the Elements Materials List

Lesson	Strategy or Activity	Materials Needed
9.1	Engagement Activity	Index Cards
9.4	Exploration Activity	Graph Paper

Multimedia Resources

You may find these additional Web-based resources helpful when teaching *Relationships Between the Elements.*

- http://dayah.com/periodic/ Interactive periodic table.
- http://www.uky.edu/Projects/Chemcomics/ Periodic table of comic books.

Possible Misconceptions

Identify: Students may think that elements exist in their elemental state in nature.

Clarify: A chemical element is a pure substance that consists of one type of atom. There are relatively few elements that exist in their elemental state. Most elements occur only in compounds with other elements.

Making the FlexBook Flexible

An important advantage of the FlexBook is the ability it gives you, the teacher, to select the chapters and lessons that you think are most important for your own classes. You should also consult the standards correlation table that follows when selecting chapters and lessons to include in the FlexBook for your classes.

TABLE 9.3: Standards Addressed by the Lessons in Relationships Between the Elements

Lesson	California Standards	SSES Standards	AAAS Standards
Lesson 9.1	1a, 1b, 1c		
Lesson 9.2			
Lesson 9.3	2e		
Lesson 9.4	1g		
Lesson 9.5			
Lesson 9.6	1f		

9.1 Families on the Periodic Table

Key Concepts

In this lesson students explore the electron configurations of families of elements.

Lesson Objectives

- Describe the patterns that exist in the electron configurations for the main group elements.
- Identify the columns in the periodic table that contain 1. the alkali metals, 2. the alkaline earth metals, 3. the halogens, and 4. the noble gases, and describe the differences between each family's electron configuration.
- Given the outermost energy level electron configuration for an element, determine its family on the periodic table.

Lesson Vocabulary

group columns of the periodic table.

period Horizontal rows of the periodic table.

alkali metals Group 1 in the periodic table *(Li, Na, K, Rb, Cs, Fr)*.

alkaline earth metals Group 2 in the periodic table *(Be, Mg, Ca, Sr, Ba, Ra)*.

noble gases Group 18 in the periodic table *(He, Ne, Ar, Kr, Xe, Rn)*.

halogens Group 17 in the periodic table *(F, Cl, Br, I, At)*.

main group elements Equivalent to the $s+p$ blocks of the periodic table, also known as "representative elements."

Strategies to Engage

- Have each student write the names and electron configurations of the first 18 elements on separate index cards. Instruct students to put the elements with the same number of valence electrons in the same column. Then ask them to move the elements that are in the same column so that the number of valence electrons increases from left to right. Next, instruct students to put the elements in the same row that contain the same number of energy levels. Then, ask them to arrange the elements so that the number of energy levels increases from top to bottom. Instruct students to compare their cards to the periodic table. Facilitate a discussion with students about the patterns that they have observed.

Strategies to Explore

- This lesson includes an introduction to several chemical families on the periodic table. Before reading, prepare less proficient readers by having students write the following on the top of separate sheets of notebook paper:

1A

2A

5A

6A

7A

8A

As they read each section have them write key points under each heading. This will give the students a quick reference and help them to organize the information. Instruct students to write a one-paragraph summary of the information they have read in each section. DI **Less Proficient Readers**

Strategies to Extend and Evaluate

- Have students create a family photo album of one of the main group elements. Instruct students to include photos of the other elements in their chosen element's family.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 9.1 Review Questions that are listed at the end of the lesson in their FlexBook.

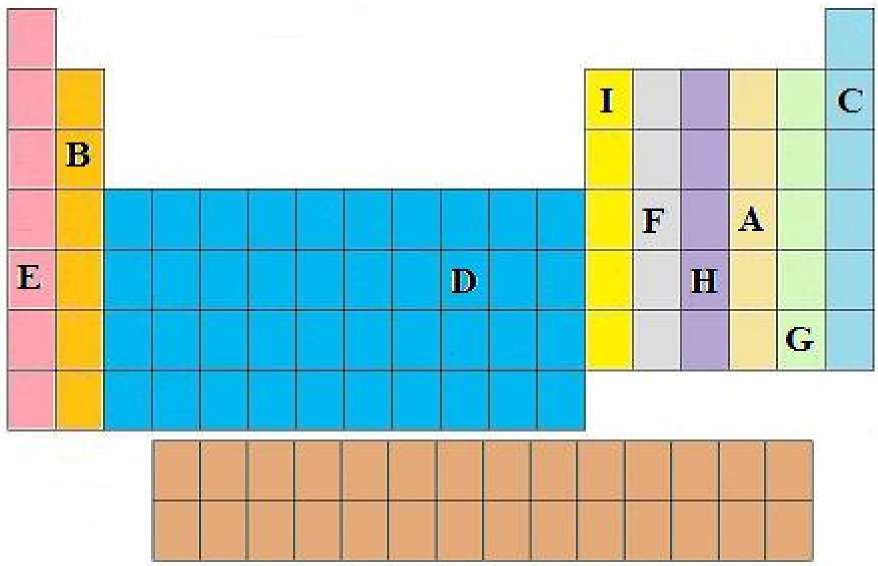
Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. If an element is said to have an outermost electronic configuration of ns^2np^3, it is in what group in the periodic table?
   a. Group 3A
   b. Group 4A
   c. Group 5A
   d. Group 7A
2. What is the general electronic configuration for the Group 8A elements? (Note: when we wish to indicate an electron configuration without specifying the exact energy level, we use the letter "n" to represent any energy level number. That is, ns^2np^3 represents any of the following; $2s^22p^3, 3s^23p^3, 4s^24p^3$, and so on.)
   a. ns^2np^6

b. ns^2np^5
c. ns^2np^1
d. ns^2

3. The group 2 elements are given what name?
   a. alkali metals
   b. alkaline earth metals
   c. halogens
   d. noble gases



4. Using the diagram above, identify:
   a. The alkali metal by giving the letter that indicates where the element would be located and write the outermost electronic configuration.
   b. The alkaline earth metal by giving the letter that indicates where the element would be located and write the outermost electronic configuration.
   c. The noble gas by giving the letter that indicates where the element would be located and write the outermost electronic configuration.
   d. The halogen by giving the letter that indicates where the element would be located and write the outermost electronic configuration.
   e. The element with an outermost electronic configuration of s^2p^3 by giving the letter that indicates where the element would be located.
   f. The element with an outermost electronic configuration of s^2p^1 by giving the letter that indicates where the element would be located.

5. In the periodic table, name the element whose outermost electronic configuration is found below. Where possible, give the name of the group.
   a. $5s^2$
   b. $4s^23d^{10}4p^1$
   c. $3s^23p^3$
   d. $5s^24d^{10}5p^2$
   e. $3s^1$
   f. $1s^2$
   g. $6s^25d^{10}6p^5$
   h. $4s^24p^4$

9.2 Electron Configurations

Key Concepts

In this lesson students explore a shortcut method for writing electron configurations.

Lesson Objectives

- Convert from orbital representation diagrams to electron configuration codes.
- Distinguish between outer energy level (valence) electrons and core electrons.
- Use the shortcut method for writing electron configuration codes for atoms and ions.

Lesson Vocabulary

orbital box diagram A diagram for drawing the electron configurations where sub-levels are shown in groups (or even in boxes) and each orbital has its own line (or box) within each sub-level.

isoelectronic Having the same electron configuration.

core electrons Electrons that occupy energy levels below the outermost energy level.

valence electrons Electrons that occupy the outer shell of the atom or ion.

Strategies to Engage

- Write the electron configuration for barium on the board. Then write its electron configuration using the shortcut method. Ask students which they would prefer to write. Inform students that, in this lesson they will learn a shortcut method of writing electron configuration. Give students the opportunity to try to figure out the shortcut method. Tell them to "stay tuned" to see if they are correct.

Strategies to Explore

- You may want to spend time exploring the effect of adding electrons to and removing electrons from a neutral atom on its charge and the fact that an atom with a charge is called an ion. Students are often confused that adding an electron to a neutral atom results in an ion with a negative charge. Remind students that electrons are negative. You can point out this fact to students by using an analogy. Tell students, "suppose you have a

group of friends and a few of those friends are negative. If you get rid of your negative friends, your group becomes more positive. If you add negative friends to your group, your group becomes more negative."

Strategies to Extend and Evaluate

- Have students play a review game called, "Two Truths and a Lie" using what they know about electron configuration. To do this, pair students, and have each pair write three statements, two of which are facts about electron configuration, and one of which is a plausible "lie." Then have each pair join with two other pairs to share what they wrote and try to guess which of the statements are "lies" and which are "truths."

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled ***Electron Configurations.*** Ask students to complete the worksheet alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 9.2 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. What is the difference between the standard electron configuration code and the electron configuration code using noble gas configurations?
2. The standard electron configuration is often called the ground state electron configuration. Why do you think this is so?
3. How is it possible to have the same number of core electrons and valence electrons for two different ions?
4. Draw the orbital representation for the electron configuration of potassium, K.
5. Write the electron configuration code for potassium, K.
6. Write the noble gas electron configuration code for potassium, K.
7. How many core electrons and valence electrons are in potassium, K?
8. Write the electron configuration for potassium, K^+.
9. Write the noble gas electron configuration code for potassium, K^+.
10. With what species is K^+ isoelectronic?
11. The electron configuration below is for which element?

$$[\text{Kr}]4d^{10}5s^25p^2$$

   a. *Sb*
   b. *Sn*
   c. *Te*
   d. *Pb*

12. The electron configuration below is for which element?

$$[\text{Xe}]5d^{10}6s^26p^4$$

a. Pb
b. Bi
c. Po
d. Tl

13. What is the noble gas electron configuration for the bromine element?

a. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 4p^5$
b. $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^5$
c. $[\text{Ar}]4s^2 4p^5$
d. $[\text{Ar}]3d^{10} 4s^2 4p^5$

14. Write the noble gas electron configuration for each of the following.

a. Al
b. N^{3-}
c. Sr^{2+}
d. Sn^{2+}
e. I

15. How many core electrons and valence electrons are in each of the following?

a. Mg
b. C
c. S
d. Kr
e. Fe

9.3 Lewis Electron Dot Diagrams

Key Concepts

In this lesson students explore electron dot diagrams.

Lesson Objectives

- Explain the meaning of an electron dot diagram.
- Draw electron dot diagrams for given elements.
- Describe the patterns of electron dot diagrams in the periodic table.

Lesson Vocabulary

Lewis Electron Dot Diagram A shorthand visual representation of the valence electrons for an element. (Lewis electron dot diagram for sodium with one valence electron: $Na^{\bullet}$

Strategies to Engage

- Have students read the review questions at the end of this section. This way, students will be familiar with the types of information that they will explore in this lesson.

Strategies to Explore

- Have students choose a period of elements (except period 1) and draw the Lewis dot structure for each element in the period. Instruct them to write a paragraph explaining any patterns they observe.

Strategies to Extend and Evaluate

- Have students create a short lesson outlining how to write Lewis dot diagrams. Encourage students to create diagrams to include in the lesson.

Lesson Worksheets

Copy and distribute the worksheet in the *CK-12 Chemistry Workbook* titled ***Electron Dot Formulas.*** Ask students to complete the worksheet alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 9.3 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. What is the advantage of the Lewis electron dot diagram?
2. What is the maximum number of dots in a Lewis's Electron dot diagram?
3. Draw the Lewis electron dot diagram for lithium.
4. Draw the Lewis electron dot diagram for calcium.
5. Draw the Lewis electron dot diagram for bromine.
6. Draw the Lewis electron dot diagram for selenium.
7. Write the Lewis electron dot diagrams for each of the following. What observations can you make based of these diagrams?
   a. Oxygen 2- ion
   b. Sulfur 2- ion
   c. Antimony
   d. Aluminum
8. Write the trend for the Lewis electron dot diagrams for Group 6A (or Group 16) by filling in the table below.

TABLE 9.4:

Element	# Valence e^-	Lewis Electron Dot Diagram
Oxygen (O)		
Sulfur (S)		
Selenium (Se)		
Tellurium (Te)		
Polonium (Po)		

9.4 Chemical Family Members Have Similar Properties

Key Concepts

In this lesson students explore trends in chemical reactivity within chemical families.

Lesson Objectives

- Explain the role of the core electrons.
- Explain the role of valence electrons in determining chemical properties.
- Explain how the chemical reactivity trend in a chemical family is related to atomic size.

Lesson Vocabulary

noble gas core When working with noble gas electronic configurations the core electrons are those housed in the noble gas symbolic notation.

Strategies to Engage

- Facilitate a discussion with students about family characteristics they share. Call on volunteers to share with the class any characteristics they may share with their family members. Tell students that in this lesson they will learn about characteristics, or properties shared by elements in the same family.

Strategies to Explore

- Have students create a graph of atomic number vs. atomic radius for elements 1-18. Facilitate a discussion with students about the patterns they observe. (Within the same period, the atomic radius decreases as you move from left to right across the period. This is because there is an increase in the nuclear charge and an increase in the number of electrons in the same energy level. Increasing the amount of nuclear charge attracts the electrons closer to the nucleus. Within the same group, the atomic radius increases from top to bottom down the group. Although there is an increase in nuclear charge, adding another principal energy level results in the valence electrons being further from the nucleus.)

Strategies to Extend and Evaluate

- Have students choose a group of elements and write a fictional story entitled, "The Day the _________ Disappeared", imagining if one day that family of elements were to disappear. Encourage students to be as creative as possible, but to include factual information about the uses of the elements in the group, and how life on Earth would be affected if that group were to disappear.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 9.4 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. What is the difference between valence electrons and core electrons?
2. Why would the valence electrons be responsible for the chemical reactivity?
3. Which of the following pairs of elements would have the greatest similarities in terms of chemical properties?
   a. Ca and Ca^{2+}
   b. Ca and Mg
   c. Ca and K
   d. Ca and Ar
4. What is the correct number of core electrons in the phosphorous atom?
   a. 3
   b. 5
   c. 10
   d. 15
5. What is the correct number of valence electrons in the iodine atom?
   a. 4
   b. 5
   c. 6
   d. 7
6. For the valence electrons in Group 6A, what conclusions can you draw about the trend in chemical reactivity?
7. For the valence electrons in Group 7A, what conclusions can you draw about the trend in chemical activity?

9.5 Transition Elements

Key Concepts

In this lesson students explore electron configurations of transition elements.

Lesson Objectives

- Define transition metals.
- Explain the relationship between transition metals and the d sublevels.
- State the periods that contain transition metals.
- Write electron configurations for some transition metals.

Lesson Vocabulary

transition metal Groups 3 through 12 or the d block of the periodic table.

Strategies to Engage

- Ask students to identify the transition metals on the periodic table. Ask them to identify some properties of transition metals. (They are shiny, hard, dense, good conductors of heat and electricity, and have high melting points.)

Strategies to Explore

- Have less proficient readers make a main ideas/details chart as they read the lesson. Instruct them to divide a sheet of paper down the middle and record the main ideas on the left side and the details for each main idea on the right side. Have students save their chart for reviewing lesson content. DI **Less Proficient Readers**

Strategies to Extend and Evaluate

- Have students research the use of transition metals in the creation of U.S. coins. Students should write a report that lists the transition metals that are found in each coin and why these metals are ideal for coins.

- Have students complete the lab ***Paramagnetism of Manganese Compounds***. This lab is located in the Supplemental Lab Book.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 9.5 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Write the electron configuration for zirconium, Zr.
2. How many valence electrons does zirconium, Zr, have?
3. Write the noble gas electronic configuration for platinum, Pt.
4. How many valence electrons does platinum, Pt, have?
5. Why do the d block elements only start in the fourth period?
6. What do copper, silver, and gold have in common as far as their electron configuration?
7. Which of these is the electron configuration for nickel?
   a. [Kr]$3d^8 4s^2$
   b. [Kr]$3d^{10}$
   c. [Ar]$3d^8 4s^2$
   d. [Ar]$3d^{10}$
8. How many d electrons are there in the electronic configuration for ruthenium?
   a. 0
   b. 6
   c. 7
   d. 17
9. Write the electron configuration for Iridium, Ir.
10. What are the valence electrons for Iridium, Ir?
11. Write the noble gas electron configuration for mercury, Hg.
12. How many valence electrons does mercury, Hg, have?

9.6 Lanthanide and Actinide Series

Key Concepts

In this lesson students explore the electron configurations of lanthanides and actinides.

Lesson Objectives

- Define the lanthanides and actinides.
- Place the lanthanides and actinides in the periodic table.
- Explain the importance of both the lanthanides and actinides.
- Write electron configurations for lanthanides and actinides.

Lesson Vocabulary

lanthanide The rare earth elements found in the first period of the f block. These elements fill up the $4f$ sublevel.

actinide The elements found in the second period of the f block. These elements fill up the $5f$ sublevel.

Strategies to Engage

- Students may wonder why the inner transition elements are often offset below the main body of the periodic table. Have them draw the block diagram located below the introduction paragraph to this lesson, but instruct them to place the lanthanides and actinides within the main body of the periodic table. Students should be able to see why the transition elements often appear below the main body of the periodic table.

Strategies to Explore

Strategies to Extend and Evaluate

- Have students choose a lanthanide or actinide of interest to them. Instruct students to research their element's properties and history and create a resume for that element. Students may use the planning page found at http://www.nytimes.com/learning/teachers/studentactivity/20090217.pdf to guide their research.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 9.6 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Why are the f block elements referred to by some as inner transition elements?
2. What do europium and americium have in common as far as their electron configuration?
3. What is the electron configuration for Berkelium?
    a. $[\text{Xe}]7s^2 5f^9$
    b. $[\text{Xe}]7s^2 5f^9 6d^1$
    c. $[\text{Rn}]7s^2 5f^9$
    d. $[\text{Rn}]7s^2 5f^9 6d^1$
4. How many f electrons are there in the electron configuration for einsteinium?
    a. 0
    b. 11
    c. 14
    d. 25
5. Write the electron configuration for Ytterbium, Yb.
6. What are the valence electrons for Ytterbium, Yb? What periods and sublevels are they in?
7. Write the noble gas electronic configuration for uranium, U.
8. What are the valence electrons for uranium, U? What periods and sublevels are they in?
9. Write the electron configurations for neptunium and then for plutonium. Now write an explanation for what seems to be happening.

CHAPTER 10 Trends on the Periodic Table TE

Chapter Outline

Trends on the Periodic Table

Outline

The chapter *Trends on the Periodic Table* consists of three lessons that explain the periodic change in atomic size and its relationship to the periodic trends for ionization energy and electron affinity.

- **Lesson 10.1 Atomic Size**
- **Lesson 10.2 Ionization Energy**
- **Lesson 10.3 Electron Affinity**

Overview

In these lessons, students will explore:

- The periodic trends in atomic size.
- The trends in ionization energy and ionic size.
- The trends in electron affinity.

Science Background Information

This information is provided for teachers who are just beginning to instruct in this subject area.

- The Chemistry of Glass

In the world of the twenty-first century, it is difficult to imagine our day-to-day existence without glass. As a transparent material with great resistance to corrosive substances, it has many uses from container duty, to its architectural impact as windows, to its role in optical devices. Its versatility places glass in an indispensable position in our array of material choices. Glass is chemically defined as a member of a group of compounds that solidify from the molten state without crystallization.

The original formulation of glass may have been based on observation of the naturally occurring opaque black glass known as obsidian, produced when volcanic lava is cooled abruptly by contact with water. Whether it was inspired or the result of a fortuitous accident, the first glasses originated in the Near East about 3000 B.C. Instructions for glassmaking have been identified on Mesopotamian cuneiform tablets. Their basic "recipe" included the same three main ingredients utilized to create modern glass formulations: formers, flux and stabilizers. A former is a material that forms the basis upon which the rest of the formulation is set; in many glasses, old and new, the former is silica, SiO_2, or sand. The flux is the substance added in a minor quantity, mainly to lower the melting temperature of the mixture. Alkalis such as soda (sodium carbonate) and potash (potassium carbonate) have long been employed in this capacity. Lastly, the stabilizer, such as lime, CaO, calcium oxide, strengthens the glass and also adds water resistance.

Various inorganic materials have been added to this classic formulation since antiquity and their incorporation has been shown to impart novel characteristics to the glass produced. One of the earliest substitutions seen in glass formulations include the addition of lead oxide as a flux material. Lead glass, which may have first been used in Han dynasty China, was shaped into artificial gemstones, and later for lead crystal stemware. Lead glass has a lower melting temperature and a reputation for brilliance and sparkle due to its high refractive index. Lead glassware is also known for its ability to "ring" when struck, that distinguishes it from ordinary silicate glass.

The addition of small amounts of various metal oxide salts to the basic glass formula produce "stained" glass, renowned in its use in churches and cathedrals. Cobalt oxide is responsible for the vivid blue coloration, red from gold salts and copper oxide imparts a brilliant green hue.

The use of borax (boric oxide, B_2O_3) in place of soda and lime was a nineteenth century innovation attributed to Otto Schott, a German glassmaker, who originally called this new material "Duran". Borosilicate glass has a higher melting temperature and a much greater thermal stability. Its modern commercial name "Pyrex", is well known both to cooking enthusiasts and laboratory scientists.

The underlying chemical rationale for the differing properties of these disparate glass formulations may be due to atomic size mismatches between the various components. Since glass is not a crystalline substance, without a regular, repeating microscopic structure, it is better represented as a disordered network with defects, or "empty spaces" in the network. The presence of atoms with variously sized atomic radii in these defect regions, can then alter the macroscopic characteristics of the glass. For example, the substitution of the smaller boron atom in place of larger alkali metals may provide more efficient silica packing, and possibly account for the enhanced thermal stability of Pyrex. Lead ions, more comparably sized to the alkali ions, present a glass product that has similar melting characteristics.

What does the future hold for new glasses? Silicate fiber optics play a vital role in modern telecommunications. Heat resistant glasses are employed on the exterior of spacecraft for protection upon re-entry into Earth's atmosphere. Smart glass windows can control the amount of incoming solar radiation, and there is research into "self-cleaning" window glass. Chicago's Sears Tower recently installed a glass observation deck flooring with load-bearing tempered glass. The material that caught the eye and imagination of humans many years ago may have many more surprising applications in store.

Pacing the Lessons

Use the table below as a guide for the time required to teach the lessons of *Trends on the Periodic Table*.

TABLE 10.1: 60 Minute Class Periods per Lesson

Lesson	Number of Class Periods
10.1 *Atomic Size*	0.5
10.2 *Ionization Energy*	0.5
10.3 *Electron Affinity*	0.5

TABLE 10.1: (continued)

Lesson	Number of Class Periods

Managing Materials

The following items are needed to teach the strategies and activities described in the Teachers Edition of the FlexBook for *Trends on the Periodic Table*.

TABLE 10.2: Trends on the Periodic Table Materials List

Lesson	Strategy or Activity	Materials Needed
10.1		
10.2		
10.3		

Multimedia Resources

You may find these additional web-based resources helpful when teaching *Trends on the Periodic Table.*

- Key information about the chemical elements. http://www.webelements.com/

Possible Misconceptions

Identify: Students may think that the elements on the periodic table always follow the repeating patterns in chemical and physical properties. Students may not know that the patterns illustrated by the periodic table are general trends, and that there are some exceptions.

Clarify: Explain to students that the periodic table, arranged in order of increasing atomic numbers of the chemical elements reveals a tendency for the chemical and physical properties of the elements to repeat in a periodic pattern. These periodic trends exist for many properties of the elements. Emphasize the use of the word "general" when describing the repeating patterns present in the periodic table.

Promote Understanding: Create an element card for each of the first 18 elements. On each card write properties of the element such as boiling point, melting point, and density. Remove or cover all copies of the periodic table in the classroom. Give one set of cards to each group of three or four students. Instruct students to create their periodic table by arranging the cards according to the properties. Students should notice that they are able to arrange most of the elements in the correct order.

Discuss: At the end of the lesson ask students, 'What is the difference between the following two statements?'

- 1. The elements on the periodic table always follow the repeating patterns in chemical and physical properties.
- 2. The elements on the periodic table have a tendency to follow repeating patterns in chemical and physical properties.

The second statement leaves room for exceptions, while the first statement does not.

Ask students: Which of the two statements is a more correct description of the periodic table of elements? Explain.

The second statement, because there are some exceptions to the periodic patterns.

Ask: How can the second statement be stated differently?

Sample answers:

- 1. The elements on the periodic table generally follow repeating patterns in chemical and physical properties.

Making the FlexBook Flexible

An important advantage of the FlexBook is the ability it gives you, the teacher, to select the chapters and lessons that you think are most important for your own classes. You should also consult the standards correlation table that follows when selecting chapters and lessons to include in the FlexBook for your classes.

TABLE 10.3: Standard Addressed by the Lessons in Trends on the Periodic Table

Lesson	California Standards	NSES Standards	AAAS Benchmarks
10.1	1c		
10.2	1c, 2g		
10.3	1c, 2g		

10.1 Atomic Size

Key Concepts

In this lesson students explore the periodic trends in atomic size.

Lesson Objectives

- Define atomic radius.
- State the boundary issue with atomic size.
- Describe measurement methods for atomic size.
- Define the shielding effect.
- Describe the factors that determine the trend of atomic size.
- Describe the general trend in atomic size for groups and for periods.
- Describe the trend of atomic radii in the rows in the periodic table.
- Describe how the trend of atomic radii works for transition metals.
- Use the general trends to predict the relative sizes of atoms.

Lesson Vocabulary

atomic size Atomic size is the distance from the nucleus to the valence shell where the valence electrons are located.

atomic radius One-half the distance between the centers of the two atoms of a homonuclear molecule.

nuclear charge The number of protons in the nucleus.

shielding effect The core electrons in an atom interfere with the attraction of the nucleus for the outermost electrons.

electron-electron repulsion The separation that occurs because electrons have the same charge.

Strategies to Engage

- Choose an element. Invite students to name three things they can predict about that element based on its position on the periodic table. Tell students that, in the next few lessons, they will learn how to predict even more information about an element based on its position on the periodic table.

Strategies to Explore

- Have groups of students come up with creative ways to act out the shielding effect, and demonstrate for the other members of the class. DI **English Language Learners**

Strategies to Extend and Evaluate

- Have students write a one-paragraph summary of this lesson. Inform students that they must include each of the vocabulary words in their summary.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 10.1 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Why is the atomic size considered to have "no definite boundary"?
2. How is atomic size measured?
    a. using a spectrophotomer
    b. using a tiny ruler (called a nano ruler)
    c. indirectly
    d. directly
3. Draw a visual representation of the atomic radii of an iodine molecule.
4. Which of the following would be smaller?
    a. In or Ga
    b. K or Cs
    c. Te or Po
5. Explain in your own words why Iodine is larger than Bromine.
6. What three factors determine the trend of atomic size going down a group?
7. What groups tend to show this trend?
8. Which of the following would have the largest atomic radii?
    a. Si
    b. C
    c. Sn
    d. Pb
9. Which of the following would have the smallest atomic radius?
    a. $2s^2$

   b. $4s^2 4p^3$
   c. $2s^2 2p^4$
   d. $4s^2$

10. Arrange the following in order of increasing atomic radii: Tl, B, Ga, Al, In.
11. Arrange the following in order of increasing atomic radii: Ge, Sn, C,
12. Which of the following would be larger?
    a. *Rb* or *Sn*
    b. *Ca* or *As*
13. Place the following in order of increasing atomic radii: Mg, Cl, S, Na.
14. Describe the atomic size trend for the rows in the periodic table.
15. Draw a visual representation of the periodic table describing the trend of atomic size.
16. Which of the following would have the largest atomic radii?
    a. *Sr*
    b. *Sn*
    c. *Rb*
    d. *In*
17. Which of the following would have the smallest atomic radii?
    a. *K*
    b. *Kr*
    c. *Ga*
    d. *Ge*
18. Place the following elements in order of increasing atomic radii: In, Ca, Mg, Sb, Xe.
19. Place the following elements in order of decreasing atomic radii: Al, Ge, Sr, Bi, Cs.
20. Knowing the trend for the rows, what would you predict to be the effect on the atomic radius if an atom were to gain an electron? Use an example in your explanation.
21. Knowing the trend for the rows, what would you predict to be the effect on the atomic radius if the atom were to lose an electron? Use an example in your explanation.

10.2 Ionization Energy

Key Concepts

In this lesson students explore the trends in ionization energy and ionic size.

Lesson Objectives

- Define ionization energy.
- Describe the trend that exists in the periodic table for ionization energy.
- Describe the ionic size trend that exists when elements lose one electron.

Lesson Vocabulary

ionization energy The energy required to remove an electron from a gaseous atom or ion: $\text{energy} + J(g) \rightarrow J^+(g) + e^-$ (first ionization energy).

effective nuclear charge The charge on the atom or ion felt by the outermost electrons (valence electrons).

Strategies to Engage

- Have students write down the lesson objectives, leaving about 5 or 6 lines of space in between. As you explore the lesson, have students write the "answer" to each objective.

Strategies to Explore

- Have groups of students come up with creative ways to act out the effective nuclear charge and demonstrate for the other members of the class. DI **English Language Learners**

Strategies to Extend and Evaluate

Have students find information on the properties of eight elements on the periodic table, one from each of the main groups. Have students write a short paragraph about each element, explaining, in terms of ionization energy, why their chosen elements might have the properties they do.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 10.2 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Define ionization energy and show an example ionization equation.
2. Draw a visual representation of the periodic table describing the trend of ionization energy.
3. Which of the following would have the largest ionization energy?
   a. Na
   b. Al
   c. H
   d. He
4. Which of the following would have the smallest ionization energy?
   a. K
   b. P
   c. S
   d. Ca
5. Place the following elements in order of increasing ionization energy: Na, O, Ca, Ne, K.
6. Place the following elements in order of decreasing ionization energy: N, Si, S, Mg, He.
7. Using experimental data, the first ionization energy for an element was found to be 600 kJ/mol. The second ionization energy for the ion formed was found to be 1800 kJ/mol. The third ionization energy for the ion formed was found to be 2700 kJ/mol. The fourth ionization energy for the ion formed was found to be 11600 kJ/mol. And finally the fifth ionization energy was found to be 15000 kJ/mol. Write the reactions for the data represented in this question. Which group does this element belong? Explain.
8. Using electron configurations and your understanding of ionization energy, which would you predict would have higher second ionization energy: Na or Mg?
9. Comparing the first ionization energy (IE_1) of calcium, Ca, and magnesium, Mg,:
   a. Ca has a higher IE_1 because its radius is smaller.
   b. Mg has a higher IE_1 because its radius is smaller.
   c. Ca has a higher IE_1 because its outer sub-shell is full.
   d. Mg has a higher IE_1 because its outer sub-shell is full.
   e. they have the same IE_1 because they have the same number of valence electrons.
10. Comparing the first ionization energy (IE_1) of beryllium, Be, and boron, B,:
   a. Be has a higher IE_1 because its radius is smaller.
   b. B has a higher IE_1 because its radius is smaller.
   c. Be has a higher IE_1 because its s sub-shell is full.
   d. B has a higher IE_1 because its s sub-shell is full.
   e. They have the same IE_1 because B has only one more electron than Be.

10.3 Electron Affinity

Key Concepts

In this lesson students explore the trends in electron affinity.

Lesson Objectives

- Define electron affinity.
- Describe the trend for electron affinity on the periodic table.

Lesson Vocabulary

electron affinity The energy input or output when an electron is added to a gaseous atom or ion. $T(g) + e^- \rightarrow T^-(g)$

Strategies to Engage

- Have students read the review questions at the end of this section. This way, students will be familiar with the types of information that they will explore in this section.

Strategies to Explore

- Outline the main concepts of the lesson as a class. Discuss the main concepts as you prepare the outline.

Strategies to Extend and Evaluate

- Have students record what they think is the main idea of each section. Have pairs of students come to a consensus on each main idea. Then, have each pair combine with another pair and again come to a consensus. Finally, have each group share their results with the class. DI **Less Proficient Readers**

Lesson Worksheets

Copy and distribute the worksheet in the *CK-12 Chemistry Workbook* titled ***Trends on the Periodic Table.*** Ask students to complete the worksheet alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 10.3 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Define electron affinity and show an example equation.
2. Choose the element in each pair that has the lower electron affinity:
   a. Li or N
   b. Na or Cl
   c. Ca or K
   d. Mg or F
3. Why is the electron affinity for calcium much higher than that of potassium?
4. Draw a visual representation of the periodic table describing the trend of electron affinity.
5. Which of the following would have the largest electron affinity?
   a. Se
   b. F
   c. Ne
   d. Br
6. Which of the following would have the smallest electron affinity?
   a. Na
   b. Ne
   c. Al
   d. Rb
7. Place the following elements in order of increasing electron affinity: Te, Br, S, K, Ar.
8. Place the following elements in order of decreasing electron affinity: S, Sn, Pb, F, Cs.
9. Describe the trend that would occur for electron affinities for elements in period 3. Are there any anomalies? Explain.
10. Comparing the electron affinity (EA) of sulfur, S, and phosphorus, P, :
    a. S has a higher EA because its radius is smaller.
    b. P has a higher EA because its radius is smaller.
    c. S has a higher EA because its p sub-shell is half full.
    d. P has a higher EA because its p sub-shell is half full.
    e. they have the same EA because they are next to each other in the periodic table.

CHAPTER 11 Ions and the Compounds They Form TE

Chapter Outline

Unit 4 Chemical Bonding and Formula Writing

Outline

This unit, *Chemical Bonding and Formula Writing,* includes five chapters that explore ionic and covalent bonds and naming and writing formulas for the resulting compounds.

- **Chapter 11 Ions and the Compounds They Form**
- **Chapter 12 Writing and Naming Ionic Formulas**
- **Chapter 13 Covalent Bonding**
- **Chapter 14 Molecular Architecture**
- **Chapter 15 The Mathematics of Compounds**

Overview

Ions and the Compounds They Form

This chapter explains the reasons for ion formation, ionic bonding, and the properties of ionic compounds.

Writing and Naming Ionic Formulas

This chapter develops the skills necessary to predict ionic charges, write ionic formulas, and name ionic compounds.

Covalent Bonding

This chapter explains the nature of the covalent bond, how and why covalent bonds form, which atoms form covalent bonds, and the nomenclature for binary covalent compounds.

Molecular Architecture

This chapter explains the formation of electronic and molecular geometries of covalent molecules including those that violate the octet rule. The chapter also develops the concept of polar molecules.

The Mathematics of Compounds

This chapter develops the skills involved in formula stoichiometry.

Ions and the Compounds They Form

Outline

The chapter *Ions and the Compounds They Form* consists of three lessons that explore the formation, structure, and properties of ionic compounds

- **Lesson 11.1 The Formation of Ions**
- **Lesson 11.2 Ionic Bonding**
- **Lesson 11.3 Properties of Ionic Compounds**

Overview

In these lessons, students will explore:

- Why some atoms form negative ions while others form positive ions.
- How electrons are transferred in the formation of ionic bonds.
- The structure and properties of ionic compounds.

Science Background Information

This information is provided for teachers who are just beginning to instruct in this subject area.

Chemical Bonding

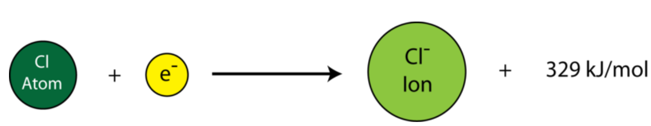
One of the primary causes of change in physical systems is the tendency toward minimum potential energy. Objects roll downhill, objects above the Earth fall, stretched rubber bands contract, objects with like charges separate, and objects with unlike charges move together. All of these changes involve a decrease in potential energy.

In many situations, the potential energy of a system increases somewhat, at first, in order to achieve a position from which the potential energy will significantly decrease. A siphon is an example of this. In a siphon, water will run up the hose as long as the final position of the water is lower in potential energy than the original position.

Ionic Bonding

In ionic bonding, the metallic atom must lose one or more electrons in order to form a bond. This loss of electrons by metallic atoms requires an input in energy. The necessary ionization energy (energy to remove an electron) must be provided to form a cation (positive ion). Suppose we use a sodium atom as an example.

An input of energy ionizes the sodium atom to a sodium ion, Na^+. In the presence of chlorine atoms, the electron can then add to a chlorine atom to form a negative chloride ion. In this process, energy is given off. The electron affinity of chlorine is $-329\ kJ/mol$. That means that adding an electron to a chlorine atom is a reduction in potential energy.



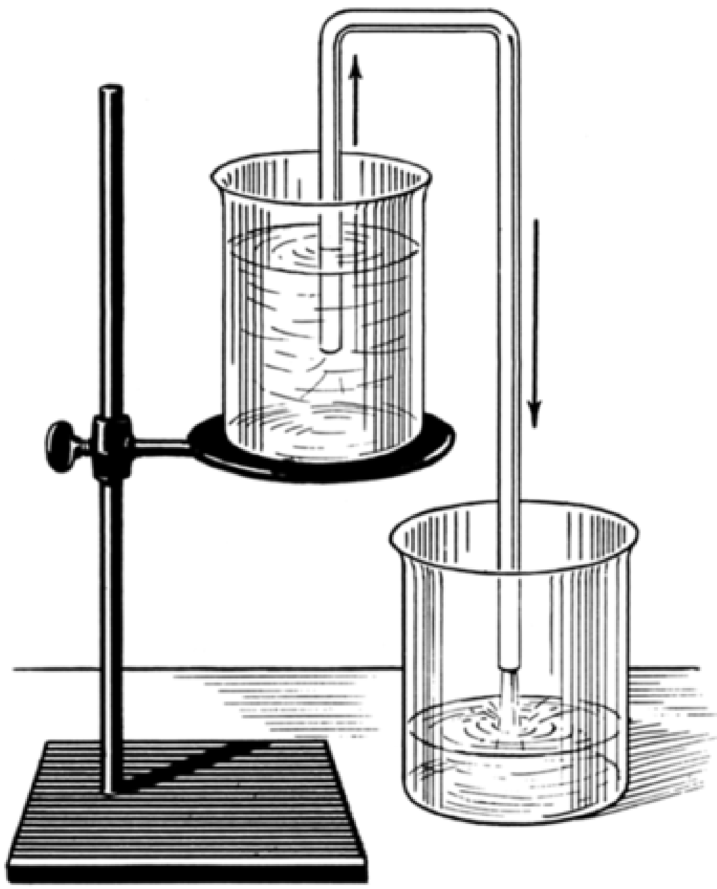
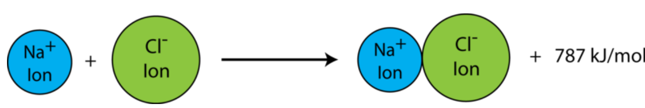


FIGURE 11.1

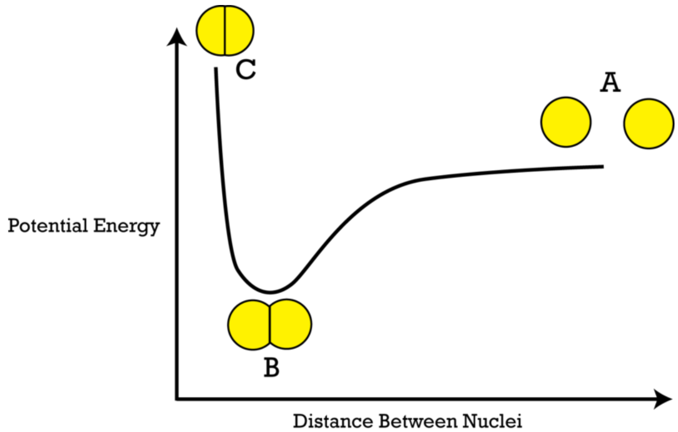
The sodium ions and the chloride ions have opposite charges, and therefore, are attracted to each other. When the ions move closer together, potential energy is again lowered. As the oppositely charged ions move together to form a crystal lattice, energy is given off. For each mole of sodium ions and chloride ions that move together to form a lattice, 787 kJ of energy are given off.



The overall process of removing an electron from a sodium atom (energy input = 496 kJ/mol), adding the electron to a chlorine atom (energy output = 329 kJ/mol), and the ions moving together in a lattice structure (energy output = 787 kJ/mol) has a net energy output of 620 kJ/mol; therefore, like a siphon, the process occurs because it has a net lowering of potential energy.

Covalent Bonding

In the case of covalent bonding, there is no electron losing, gaining, or transferring. In covalent bonds, the bonding electrons are shared. The potential energy lowering that occurs in covalent bonding can be represented by showing the relationship between the potential energy of the system and the distance between the nuclei of the bonding atoms.



When individual non-metallic atoms are completely separated, there is an attraction (electron affinity) between the electrons of each atom and the protons in the nucleus of the other atom. This attraction and the distance between the atoms cause potential energy to exist. As the atoms move closer together, the potential energy of the system decreases because the distance between the attracting objects is becoming less.

As the distance between the two nuclei decreases, the potential energy becomes less and less. Since these are non-metallic atoms, their electron outer energy levels are not full and this allows the electron clouds of the two atoms to overlap, and the atoms may continue to move closer together. When the atoms reach position B in the diagram, the potential energy of the system is at its lowest possible for these two atoms. If the nuclei continue to move closer together, the potential energy increases dramatically due to the repulsion of the positively charged nuclei. In position A, the nuclei are too close together so potential energy is high. In position C, the atoms are too far apart so the potential energy is high. In position B, the atoms are the proper distance apart for lowest potential energy. Since these two atoms reach lowest potential energy in a position where their electron clouds are overlapped, they will be bonded. The distance between the nuclei in position B is known as the bond length. If the atoms attempt to move either closer together or farther apart, the potential energy increases. The tendency toward minimum potential energy causes these atoms to remain in the bonded position.

Pacing the Lessons

Use the table below as a guide for the time required to teach the lessons of *Ions and the Compounds They Form*.

TABLE 11.1: 60 Minute Class Periods per Lesson

Lesson	Number of Class Periods
11.1 *The Formation of Ions*	0.5
11.2 *Ionic Bonding*	0.5
11.3 *Properties of Ionic Compounds*	0.5

Managing Materials

The following items are needed to teach the strategies and activities described in the Teachers Edition of the FlexBook for *Ions and the Compounds They Form*.

TABLE 11.2: Ions and the Compounds They Form Materials List

Lesson	Strategy or Activity	Materials Needed
11.1	Possible Misconceptions	Iron filings, powdered sulfur, Petri dish, Bunsen burner, magnet, metal rod
11.2	Exploration Activity	Small objects such as gum drops, beans, or paper clips.
11.3	Engagement Activity	Table salt, sugar, 2 beakers, conductivity tester

Multimedia Resources

You may find these additional web based resources helpful when teaching *Ions and the Compounds They Form*:

- Printable periodic tables for handouts: http://science.widener.edu/ svanbram/ptable.html
- Fill-in-the-blank worksheet generator: http://www.theteacherscorner.net/printable-worksheets/make-your-own/fill-in-the-blank/
- Interactive lesson on chemical bonding: http://www.visionlearning.com/library/module_viewer.php?mid=55
- Interactive lesson on ionic bonding: http://www.teachersdomain.org/resource/lsps07.sci.phys.matter.ionicbonding/

Possible Misconceptions

Identify: Students may think that ionic compounds are formed spontaneously any time a metal and a nonmetal come into contact with each other.

Clarify: When a metal and a nonmetal come into contact with each other, they can sometimes physically combine to produce a mixture.

Promote Understanding: Combine small amounts of sulfur and iron fillings in a Petri dish to produce a mixture. Run a magnet on the outside of the Petri dish to separate the iron from the sulfur without the sulfur becoming stuck to the magnet. Explain to students that the mixing of the iron and sulfur was not a chemical reaction because no new substances were formed. Re-combine the two elements into a glass dish or beaker in a fume hood. Use glassware that you don't mind ruining. Use a Bunsen burner to heat a metal rod and place it into the iron/sulfur mixture. The mixture will start to glow, and the elements will react to form iron(II) sulfide, a compound. The iron sulfide may become fused to the metal rod. Gases that are both toxic and corrosive could be produced, so perform this demonstration in a well-ventilated area. Explain to students that added heat caused a chemical reaction between the iron and the sulfur to occur. In this reaction a new substance -iron(II) sulfide- was formed.

Making the FlexBook Flexible

An important advantage of the FlexBook is the ability it gives you, the teacher, to select the chapters and lessons that you think are most important for your own classes. You should also consult the standards correlation table that follows when selecting chapters and lessons to include in the FlexBook for your classes.

TABLE 11.3: Standard Addressed by the Lessons in Ions and the Compounds They Form

Lesson	California Standards	NSES Standards	AAAS Benchmarks
11.1	2c		
11.2	2a, 2h		
11.3	2c		

11.1 The Formation of Ions

Key Concepts

In this lesson students explore why some atoms form negative while others form positive ions.

Lesson Objectives

- The students will define an ion.
- The students will identify the atoms most likely to form positive ions, and the atoms most likely to form negative ions.
- The students will explain why atoms form ions.
- The students will predict the charge on ions from the electron affinity, ionization energies, and electron configuration of the atom.

Lesson Vocabulary

ion An atom or group of atoms with an excess positive or negative charge.

cation a positive ion.

anion a negative ion.

Strategies to Engage

- On the board or chart paper, write the octet rule. (In reactions, atoms tend to lose, gain, or share electrons in order to have 8 valence electrons.) Have students write the electron configurations of each of the representative elements in period 3. Ask students to use the electron configurations and the octet rule to try to predict the number of electrons each element will gain or lose in chemical reactions. At the end of the lesson, have students see if their original answers were correct.

Strategies to Explore

- Hand each students a copy of the periodic table. As you explore this lesson, have students write the most probable ionic charges of the elements above each group of representative elements.

- Point out the division of the periodic table into metals and non-metals from the student text. Explain to students that metallic atoms tend to lose their electrons to form positively charged cations, and nonmetallic atoms tend to gain electrons to form negatively charged anions.
- Play a game with the students. Write the names of 20 elements on chart paper or the board. Have students use the periodic table to write the most probable ionic charge for each element. The student who completes the list in the fastest amount of time wins!
- Students often struggle with remembering that cations are positive and anions are negative. An easy way to distinguish and remember them is to look at the words themselves. Point out that the word "anion" can be read "a-negative-ion". The word cation can be read "c-a-positive ion".
- Have students model the creation of ions of common elements such as fluorine, magnesium, sulfur, and potassium using small objects such as gum drops, beans, or paper clips. Instruct students to use the available objects to show that positive ions are formed by the loss of electrons, and negative ions are formed by the gain of electrons. DI **English Language Learners**

Strategies to Extend and Evaluate

- Have each student write five fill-in-the-blank statements, with the blank at the end of the sentence about key concepts explored in this lesson. Have students exchange papers with another student who will try to complete the sentence by filling in the blank. Have them hand the papers back to the original student who will assign a grade. Encourage students to discuss any incorrect answers. Students can also generate fill-in-the-blank worksheets at: http://www.theteacherscorner.net/printable-worksheets/make-your-own/fill-in-the-blank/
- Have each student choose a different element and write down as much information they can about the element based on its position on the periodic table. Students should be prepared to share their information with the rest of the class. Possible information includes: atomic number; whether the element is a metal, nonmetal, or metalloid; most probable ionic charge; number of valence electrons; energy level of its valence electrons; etc.

Lesson Worksheets

Copy and distribute the worksheet in the *CK-12 Chemistry Workbook* titled ***Ion Formation Worksheet.*** Ask students to complete the worksheet alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 11.1 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Define an ion.
2. In general, how does the ionization energy of metal compare to the ionization energy of a non-metal?
3. Will an iron atom form a positive or negative ion? Why?
4. Will a bromine atom form a positive or negative ion? Why?
5. How is the number of valence electrons of a metal atom related to the charge on the ion the metal will form?
6. How is the number of valence electrons of a non-metal related to the charge on the ion the non-metal will form?

7. If carbon were to behave like a metal and give up electrons, how many electrons would it give up?
8. How many electrons are in a typical sodium ion?

11.2 Ionic Bonding

Key Concepts

In this lesson students explore how electrons are involved in the formation of ionic bonds.

Lesson Objectives

- The students will describe how atoms form an ionic bond.
- The students will state why, in terms of energy, atoms form ionic bonds.
- The students will state the octet rule.
- Given the symbol of a representative element, the students will indicate the most likely number of electrons the atom will gain or lose.
- Given the electron configuration of a representative element, the students will indicate the most likely number of electrons the atom will gain or lose.
- Given the successive ionization energies of a metallic atom, the students will indicate the most likely number of electrons the atom will lose during ionic bond formation.

Lesson Vocabulary

ionic bond A bond between ions resulting from the transfer of electrons from one of the bonding atoms to the other and the resulting electrostatic attraction between the ions.

electrostatic attraction The force of attraction between opposite electric charges.

Strategies to Engage

- Review negative and positive ions and how they are formed. Students should recall that cations are formed from the loss of electrons by a neutral atom. Inform students that, in this lesson, they will learn what happens to these "lost" electrons". Students should also recall that anions are formed from the gain of electrons by a neutral atom. Inform students that, in this lesson, they will learn where these electrons come from.

Strategies to Explore

- Have students model the creation of ionic compounds such as sodium chloride, calcium fluorine, magnesium sulfide, and lithium oxide using small objects such as gum drops, beans, or paper clips. Instruct students to use the available objects to show the transfer of electrons in each compound. DI **English Language Learners**

Strategies to Extend and Evaluate

- Have students write a creative personal ad for a representative element looking for a mate. Instruct students to include a picture and description of the element as well as a description of what the element is looking for in a "mate".

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 11.2 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. What takes place during the formation of an ionic bond?
2. What effect does the transfer of electrons have on the nuclei of the atoms involved?
3. Explain why chlorine is a small atom and tends to take on an extra electron but argon is an even smaller atom and does not tend to take on electrons.
4. If an atom had the following successive ionization energies, to which family would it belong? Why did you chose this family?

$$1^{st} \text{ ionization energy} = 75 \text{ kJ/mol}$$
$$2^{nd} \text{ ionization energy} = 125 \text{ kJ/mol}$$
$$3^{rd} \text{ ionization energy} = 1225 \text{ kJ/mol}$$
$$4^{th} \text{ ionization energy} = 1750 \text{ kJ/mol}$$

11.3 Properties of Ionic Compounds

Key Concepts

In this lesson students explore the structure and properties of ionic compounds.

Lesson Objectives

- The student will give a short, generic description of a lattice structure.
- The student will identify distinctive properties of ionic compounds.

Lesson Vocabulary

Crystal lattice A systematic, symmetrical network of atoms forming an ionic solid.

Strategies to Engage

- Sprinkle a small amount of table salt from an unlabeled container into a beaker of water. Place the electrodes of a conductivity tester into the solution and plug it in. The bulb will glow. Sprinkle a small amount of sugar from an unlabeled container into a beaker of water. Place the electrodes of the conductivity tester into the solution. The bulb will not glow. Allow students the opportunity to offer a possible explanation. Explain to the students that the first compound was table salt, an ionic compound and the second substance was sugar, which is not ionic. Inform students as you explore this lesson they will be able to explain why the ionic compound caused the bulb to glow.

Strategies to Explore

- Have students write down the lesson objectives leaving five to ten lines of space in between. As you explore the lesson, encourage students to write the "answer" to each objective.

Strategies to Extend and Evaluate

- Have each student choose an ionic compound on which to research and report. Ask students to research information such as where the compound occurs naturally, what it is used for, and its properties. Students should be prepared to share their findings with the class.

- Ask students to look at Figure 11.4. Have them write a paragraph to describing the illustration in their own words.
- Have students read the ***Softening Hard Water*** extra reading located in the Supplemental Workbook.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 11.3 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Hydrogen gas is not an acid and chlorine gas is not an acid but when hydrogen and chlorine combine to form hydrogen chloride, the compound is an acid. How do you explain that?
2. Why do we not refer to molecules of sodium chloride?
3. Which is larger, a fluorine atom or a fluoride ion?

11.4 References

1. . http://commons.wikipedia.org/wiki/File:Siphon.png,. Scott Foresman grants anyone the right to use this work for any purpose, without any conditions.

Chapter 12 Writing and Naming Ionic Formulas TE

Chapter Outline

Writing and Naming Ionic Formulas

Outline

The chapter *Writing and Naming Ionic Formulas* consists of two lessons that develop the skills involved in predicting ionic charge, writing ionic formulas, and naming ionic compounds.

- **Lesson 12.1 Predicting Formulas of Ionic Compounds**
- **Lesson 12.2 Inorganic Nomenclature**

Overview

In these lessons, students will learn how to:

- Write formulas for ionic compounds.
- Name ionic compounds.

Science Background Information

This information is provided for teachers who are just beginning to instruct in this subject area.

- The History of Chemical Symbols

The one or two letter shorthand used to represent chemical elements is a familiar feature in modern science. The tradition of using symbols to represent elements is quite ancient. Long before those interested in studying the composition and behavior of matter were known as chemists, mystical practitioners of alchemy devised symbols often to obfuscate their experimentation, and to cloak their work in secrecy. Their coded imagery drew inspiration from astrology, as well as ancient writing systems like the hieroglyphs. Alchemists linked certain metals with celestial bodies to describe their behavior, such as the connection between the rapidly moving planet Mercury and the metallic liquid quicksilver.

As chemistry became an experimental science and new methods produced scores of newly discovered elements, the need for a shorthand technique to describe chemical changes became apparent. One of the first chemists to attempt

to introduce a symbolic system for identifying the elements was John Dalton, known for his relative mass scale of the atomic weights. His symbols, introduced in 1808 in his "New System of Chemical Philosophy", consisted mainly of circles, some with inscribed alphabetic letters and others with dots or lines within the circles. Compounds were written as combinations of circles representing the constituent atoms. His system did not lend itself to ready memorization and did not catch on with his contemporaries.

Our modern method of using one or two letter shorthand for the elements was devised in 1813 by Jöns Jakob Berzelius, citing the ease of implementation, particularly for typesetters. Due to the common employment of Latin terminology in scientific communication, Berzelius suggested using the first or first two letters of the element's Latin name as the symbol for that atom. In the case of confusion or duplication of the letters, exceptions includes the use of Hg (hydrargyrum for Mercury and plumbum for lead). Some modern modifications have been introduced for new elements, especially those named in honor of famous scientists. Berzelius is also responsible for the use of subscripts in a chemical formula to designate the ratio of atoms.

Pacing the Lessons

Use the table below as a guide for the time required to teach the lessons of *Writing and Naming Ionic Formulas*.

TABLE 12.1: 60 Minute Class Periods per Lesson

Lesson	Number of Class Periods
12.1 *Predicting Formulas of Ionic Compounds*	1.0
12.2 *Ionic Nomenclature*	1.0

Managing Materials

The following items are needed to teach the strategies and activities described in the Teachers Edition of the FlexBook for *Writing and Naming Ionic Formulas*.

TABLE 12.2: Writing and Naming Ionic Formulas Materials List

Lesson	Strategy or Activity	Materials Needed
12.1	Exploration Activity	Index cards
12.2	Exploration Activity	Index cards

Multimedia Resources

You may find these additional internet resources helpful when teaching *Writing and Naming Ionic Formulas*:

- Writing the formulas of ionic compounds flowchart: http://www.phs.princeton.k12.oh.us/departments/Science/ldusch/honorspdfs/namingchpt5/Flowcharts.pdf

Possible Misconceptions

Making the FlexBook Flexible

An important advantage of the FlexBook is the ability it gives you, the teacher, to select the chapters and lessons that you think are most important for your own classes. You should also consult the standards correlation table that follows when selecting chapters and lessons to include in the FlexBook for your classes.

TABLE 12.3: **Standard Addressed by the Lessons in Writing and Naming Ionic Formulas**

Lesson	California Standards	NSES Standards	AAAS Benchmarks
12.1			
12.2			

- Apparently, the California standards do not list writing formulas or naming compounds as a standard.

12.1 Predicting Formulas of Ionic Compounds

Key Concepts

In this lesson students learn how to write formulas for ionic compounds.

Lesson Objectives

- Given the elements to be combined, the student will write correct formulas for binary ionic compounds, compounds containing metals with variable oxidation numbers, and compounds containing polyatomic ions.

Lesson Vocabulary

oxidation number The charge or apparent charge that an atom in a compound or ion would have if all the electrons in its bonds belonged entirely to the more electropositive atom.

polyatomic ion An electrically charged species formed by covalent bonding of atoms of two or more different elements, usually non-metals.

Strategies to Engage

- Review with students how to determine the number of valence electrons and most probable ionic charge for representative elements based on their position on the periodic table. Explain to students that it is also possible to predict the formulas that result from the combination of elements based on their positions on the periodic table. Write Ca and Cl on the board. Ask students to try to predict the formula for the compound that would result from the combination of elements. Inform students that in this lesson they will learn how to do just that.

Strategies to Explore

- Give each student three index cards. Have each student label one index card with each of the following: binary ionic compounds, compounds containing metals with variable oxidation numbers, and compounds containing polyatomic ions. As you explore this lesson, have students write key points under each heading. This will give students a quick reference and help them to organize the information.
- Have students create flash cards with the name of the formula of a polyatomic ion on one side and its name on the other side. Encourage students to have friends and family members quiz them until they have memorized the ten most common polyatomic ions.

Strategies to Extend and Evaluate

- Have students create a short lesson on how to write formulas for ionic compounds. Tell students to include instructions on how to write formulas for: binary ionic compounds, compounds containing metals with variable oxidation numbers, and compounds containing polyatomic ions.
- Have students organize the information explored in this lesson into a flowchart that can be used to write the formula of an ionic compound given the atoms or polyatomic ions involved. An example is shown at: http://www.phs.princeton.k12.oh.us/departments/Science/ldusch/honorspdfs/namingchpt5/Flowcharts.pdf

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled ***Formula Writing.*** Ask students to complete the worksheet alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 12.1 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Fill in the chart by writing formulas for the compounds that might form between the ions in the columns and rows. Some of these compounds don't exist but you can still write formulas for them.

TABLE 12.4:

	Na^{+}	Ca^{2+}	Fe^{3+}	$NH_4{}^{+}$	Sn^{4+}
$NO_3{}^{-}$					
$SO_4{}^{2-}$					
Cl^{-}					
S^{2-}					
$PO_4{}^{3-}$					
OH^{-}					
$Cr_2O_7{}^{2-}$					
$CO_3{}^{2-}$					

12.2 Inorganic Nomenclature

Key Concepts

In this lesson students explore naming ionic compounds.

Lesson Objectives

- Given the formula for a binary ionic compound, a compound containing metals with variable oxidation numbers, or a compound containing polyatomic ions, the students will be able to name it.
- Given the name for a binary ionic compound, a compound containing metals with variable oxidation numbers, or a compound containing polyatomic ions, the students will write the correct formula for it.

Lesson Vocabulary

anion An ion with a negative charge.

cation An ion with a positive charge.

chemical nomenclature The system for naming chemical compounds.

ionic bond The electrostatic attraction between ions of opposite charge.

polyatomic ion A group of atoms bonded to each other covalently but possessing an overall charge.

Strategies to Engage

- Write the following chemical formulas on the board: $NaCl$, K_2SO_4, and Fe_2O_3. Asks students if they can correctly state the name of each compound. Ask any student who is able to provide the correct answer to explain how they were able to correctly name the compounds. Inform students that in this lesson they will learn how to name ionic compounds.

Strategies to Explore

- Give each student three index cards. Have each student label one index card with each of the following: binary ionic compounds, compounds containing metals with variable oxidation numbers, and compounds containing

polyatomic ions. As you explore this lesson, have students write key points under each heading. This will give students a quick reference and help them to organize the information.

Strategies to Extend and Evaluate

- Have students create a short lesson on how to name ionic compounds. Tell students to include instructions on how to name: binary ionic compounds, compounds containing metals with variable oxidation numbers, and compounds containing polyatomic ions.
- Have students organize the information explored in this lesson into a flowchart that can be used to name an ionic compound given the chemical formula. An example is shown at: http://www.phs.princeton.k12.oh.us/departments/Science/ldusch/honorspdfs/namingchpt5/Flowcharts.pdf

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled ***Inorganic Nomenclature.*** Ask students to complete the worksheet alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 12.2 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Name the following compounds. CaF_2 __________________
   $(NH_4)_2CrO_4$ __________________
   K_2CO_3 __________________
   $NaCl$ __________________
   PbO __________________
   $CuSO_4$ __________________
   H_2CO_3 __________________
   $Ca(NO_3)_2$ __________________
   $Mg(OH)_2$ __________________
   SnO_2 __________________
2. Write the formulas from the names of the following compounds. Sodium carbonate __________________
   Calcium hydroxide __________________
   Iron (III) nitrate __________________
   Magnesium oxide __________________
   Aluminum sulfide __________________
   Copper (I) dichromate __________________
   Ammonium sulfate __________________
   Iron (II) phosphate __________________
   Hydrogen nitride __________________
   Lead (IV) sulfate __________________

Chapter 13 Covalent Bonding TE

Chapter Outline

Covalent Bonding

Outline

The chapter *Covalent Bonding* consists of three lessons that explain the nature of the covalent bond, which atoms form covalent bonds, and the nomenclature rules for covalent compounds.

- **Lesson 13.1 The Covalent Bond**
- **Lesson 13.2 Atoms that Form Covalent Bonds**
- **Lesson 13.3 Naming Covalent Compounds**

Overview

In these lessons, students will learn:

- How and why covalent bonds form.
- How to draw Lewis structures of molecules.
- How to apply the IUPAC nomenclature system to name binary covalent compounds.

Science Background Information

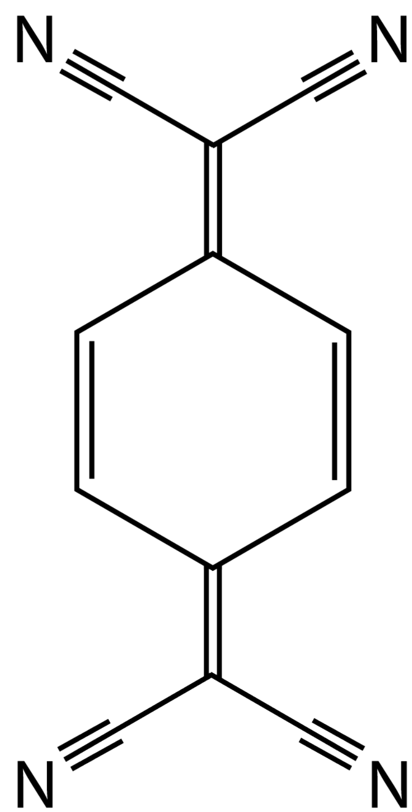
This information is provided for teachers who are just beginning to instruct in this subject area.

- **Organic Conductors**

Metallurgy was one of the first applied sciences to be mastered by humankind, and successive generations have garnered increasing expertise in processing metal ores into many different versatile and valuable substances. Despite the overwhelming dependence of modern technology on metals and their applications, new supplies of many different common metals have become increasingly more difficult to locate and procure. Many metals once considered plentiful are now deemed semi-precious, and the possibility exists that we may be restricted to the reserves on hand in the future.

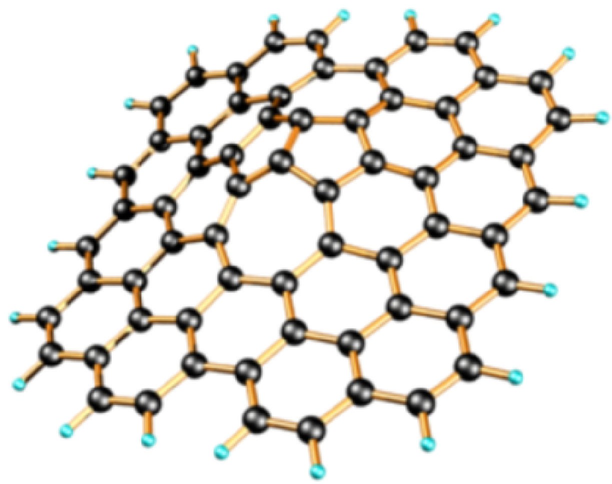
The properties that make metals so valuable, such as their electrical and heat conductivity, malleability, hardness, and density are difficult to replicate in other materials. One attempt to retain the conductive characteristics of metals in more readily available materials, is the development of organic conductors. Although most organic molecules are considered to be insulators, organic materials have been developed to produce semiconductors as well as truly conductive systems.

The first organic conductors were constructed as charge transfer complexes; these systems consisted of two molecules with one acting as an electron donor and the other an electron acceptor. For example, tetracyanoquinodimethane (TCNQ) was first identified in 1962. As its structural formula indicates, TCNQ contains alternating single, double and triple bonds, and this structure readily accepts electrons while



resulting in reallocation of the pi bonding electrons into new bonding arrangements. Several TCNQ complexes with a variety of electron donors, with high conductivities even into temperature ranges when the salt complexes melted.

Organic conductors are compelling research targets due to the vast availability of the raw materials used to prepare them, and new research suggests the possibility of producing conductive biomaterials for medical applications. The graphene molecule has already been demonstrated to form attachments with nerve cells which display electrical conductance.



Pacing the Lessons

Use the table below as a guide for the time required to teach the lessons of *Covalent Bonding*.

TABLE 13.1: 60 Minute Class Periods per Lesson

Lesson	Number of Class Periods
13.1 *The Covalent Bond*	1.0
13.2 *Atoms that Form Covalent Bonds*	1.5
13.3 *Naming Covalent Compounds*	1.0

Managing Materials

The following materials are needed to teach the strategies and activities described in the Teachers Edition of the FlexBook for *Covalent Bonding*.

TABLE 13.2: Covalent Bonding Materials List

Lesson	Strategy or Activity	Materials Needed
13.1		
13.2		
13.3		

Multimedia Resources

You may find these additional internet resources helpful when teaching *Covalent Bonding*:

Possible Misconceptions

Identify: Students may think a bond must be either ionic or covalent.

Clarify: There is a continuum of ionic and covalent character that can be assigned to a bond. In other words chemical bonds can have characteristics along a continuum from an equal sharing of electrons to a complete transfer of electrons. If the bond involves two of the same atom (A-A), then the bond must be 100% covalent because neither atom has the ability to attract the electron pair more strongly than the other. However if the bond involves different atoms (A-B), the bond will have mixed covalent and ionic character.

Making the FlexBook Flexible

An important advantage of the FlexBook is the ability it gives you, the teacher, to select the chapters and lessons that you think are most important for your own classes. You should also consult the standards correlation table that follows when selecting chapters and lessons to include in the FlexBook for your classes.

TABLE 13.3: Standard Addressed by the Lessons in Covalent Bonding

Lesson	California Standards	NSES Standards	AAAS Benchmarks
13.1	2a		
13.2	2b, 2e		
13.3	2b		

13.1 The Covalent Bond

Key Concepts

In this lesson students explore covalent bond formation.

Lesson Objectives

- The students will describe how covalent bonds form.
- The students will explain the difference between ionic and covalent bond formation and structure.
- The students will state the relationship between molecular stability and bond strength.

Lesson Vocabulary

covalent bond A type of chemical bond where two atoms are connected to each other by the sharing of two or more electrons in overlapped orbitals.

covalent bond strength The strength of a covalent bond is measured by the amount of energy required to break the bond.

Strategies to Engage

- Review ionic bonding with students. Remind students that in ionic bonding, electrons leave metallic atoms and enter nonmetallic atoms. This complete transfer of electrons changes both of the atoms into ions. Inform students that, in this lesson, they will explore the bonding that occurs between nonmetallic atoms. Give students an opportunity to try to figure out how this type of bonding occurs.

Strategies to Explore

- Instruct students to summarize the information in the section, *Molecular Stability* into a table, concept map, or other diagram.

Strategies to Extend and Evaluate

- Outline the main concepts of the lesson as a class. Discuss the main concepts as you prepare the outline.

- Read each statement in the lesson summary. Have students indicate whether or not they understand each statement by using thumbs up or thumbs down to show "Yes" or "No". Whenever a student uses a thumbs down to show "No", use this as an opportunity to review this concept with the class. DI **English Language Learners**

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 13.1 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Describe the characteristics of two atoms that would be expected to form an ionic bond.
2. Describe the characteristics of two atoms that would be expected to form a covalent bond.
3. If an atom had a very high bond energy, would you expect it to be stable or unstable?
4. When gaseous potassium ions and gaseous fluoride ions join together to form a crystal lattice, the amount of energy released is 821 kJ/mol. When gaseous potassium ions and gaseous chloride ions join together to form a crystal lattice, the amount of energy released is 715 kJ/mol. Which is the stronger bond, KF or KCl? It these two compounds were increasingly heated, which compound would break apart at the lower temperature?

13.2 Atoms that Form Covalent Bonds

Key Concepts

In this lesson students explore Lewis structures of molecules.

Lesson Objectives

- The students will identify pairs of atoms that will form covalent bonds.
- The students will draw Lewis structures for simple covalent molecules.
- The students will identify sigma and pi bonds in a Lewis structure.

Lesson Vocabulary

covalent bond A type of bond in which electrons are shared by atoms.

diatomic molecule A molecule containing exactly two atoms.

double bond A bond in which two pairs of electrons are shared.

triple bond A bond in which three pairs of electrons are shared.

sigma bond A covalent bond in which the electron pair is shared in an area centered on a line running between the atoms.

pi bond A covalent bond in which *p* orbitals share an electron pair occupying the space above and below the line joining the atoms.

Strategies to Engage

- Have students read the review questions at the end of this section. This way, students will be familiar with the types of information that they will explore in this section.

Strategies to Explore

Strategies to Extend and Evaluate

- Have students create a picture dictionary for the six vocabulary words shown. They should draw their own illustrations to help explain the meaning of each term, and write the definition under or beside the picture. DI **English Language Learners**
- Have students choose the ten sentences from the text that most closely represent the main ideas of this lesson. Have them turn these sentences into a one-two paragraph summary of this lesson.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 13.2 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Which of the following compounds would you expect to be ionically bonded and which covalently bonded?

TABLE 13.4:

Compound	Ionic or Covalent
CS_2	
K_2S	
FeF_3	
PF_3	
BF_3	
AlF_3	
BaS	

2. How many sigma bonds and how many pi bonds are present in a triple bond?
3. Draw the Lewis structure for CCl_4.
4. Draw the Lewis structure for SO_2.

13.3 Naming Covalent Compounds

Key Concepts

In this lesson students learn the IUPAC nomenclature system for naming binary covalent compounds.

Lesson Objectives

- The students will name binary covalent compounds using the IUPAC nomenclature system.
- The students will provide formulas for binary covalent compounds given the IUPAC name.

Lesson Vocabulary

Strategies to Engage

- Write the following chemical formulas on the board: CO_2, N_2O_3, and PCl_3. Asks students if they can correctly state the name of each compound. Ask any student who is able to provide the correct answer, to explain how they were able to correctly name the compounds. Inform students that in this lesson they will learn how to name covalent compounds.

Strategies to Explore

Strategies to Extend and Evaluate

- Have students create a short lesson on how to name covalent compounds.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 13.3 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Name the compound CO.
2. Name the compound PCl_3.
3. Name the compound PCl_5.
4. Name the compound N_2O_3.
5. Name the compound BCl_3.
6. Name the compound SF_4.
7. Name the compound Cl_2O.
8. Write the formula for the compound sulfur trioxide.
9. Write the formula for the compound dinitrogen tetrafluoride.
10. Write the formula for the compound oxygen difluoride.
11. Write the formula for the compound dinitrogen pentoxide.
12. Write the formula for the compound sulfur hexafluoride.
13. Write the formula for the compound tetraphosphorus decaoxide.

CHAPTER 14 Molecular Architecture TE

Chapter Outline

Molecular Architecture

Outline

The chapter *Molecular Architecture* consists of five lessons that cover the electronic and molecular geometries of covalent molecules including those that break the octet rule and the theories involved in explaining them. The chapter also develops the concept of polar molecules.

- **Lesson 14.1 Types of Bonds that Form Between Atoms**
- **Lesson 14.2 The Covalent Molecules of Family 2A-8A**
- **Lesson 14.3 Resonance**
- **Lesson 14.4 Electronic and Molecular Geometry**
- **Lesson 14.5 Molecular Polarity**

Overview

In these lessons, students will explore:

- The relationship between electronegativity and bond type.
- Hybridization in various molecules.
- Resonance structures of covalent molecules.
- The use of VSEPR theory in determining the molecular geometry of covalent compounds.
- How to determine molecular polarity

Science Background Information

This information is provided for teachers who are just beginning to instruct in this subject area.

- ***Chelates***

The expected shapes of molecules containing non-metal atoms can be predicted from Valence Shell Electron Pair Repulsion (VSEPR) Theory. The basis of this theory dictates that the optimal shape of the molecule maximizes the spatial distance between groups situated around a central atom.

Metals also, may have groups oriented around them utilizing the same premises for assigning their shape. In the case of metal ions, the attached groups are usually referred to as ligands. When one ligand is attached to more than one site in the coordination sphere of the central metal, this is an example of a group known as a chelate. The term chelate comes from the Greek word "chele" meaning the claw, such as that of a crab or lobster. The ready attachment of these multidentate groups has been employed to extract the metal ion in certain situations, such as in what is known as chelation therapy. This technique is used to remove certain undesirable or toxic metal ions, such as lead or mercury ions, from the body in cases of heavy metal poisoning.

The first use of chelating agents was between the world wars as an antidote to the arsenic-based poisonous gas, Lewisite, used on the battlefields of World War I. With what became known as British anti-Lewisite (BAL), a sulfur–based chelation agent was successfully applied to treat the gassing victims. In addition, the application of a chelate can be used to sequester metal ions such as radioactive thorium or plutonium for waste stream remediation.

Chelates have also been used to stabilize metal ions and in some cases, improve their solubility as well. Gadolinium ions are desirable paramagnetic agents for use as contrast agents in Magnetic Resonance Imaging, although the metal ions themselves have considerable toxicity. The use of DTPA (Diethylene triamine pentaacetic acid), has proved to be an effective agent for the enhanced solubility, improved biodistribution, but most importantly, superior stability in vivo. Gd-DTPA contrast agents were approved for use in human MRI scans in 1988.

Chelates have also been used in metalworking applications to control the availability of the metal ion. In many cases, chelates are used in place of other more toxic ligands, such as the cyanide ion.

Metal chelates are also employed in agricultural applications to provide improved interaction of metal ions with soil components. Also for better migration of the metal ion and therefore better distribution, particularly for those metal ions with important roles as macronutrients and micronutrients.

One possible application of the use of chelates in medical treatment may be their use in the arteriosclerosis therapy. Research in progress utilizes chelates to sequester the calcium ions in arterial plaques. As calcium ions may serve as the binders that keep these plaques intact, the exploitation of the chelate effect may prove to be a key breakthrough in improving the longterm health of cardiac patients.

Pacing the Lessons

Use the table below as a guide for the time required to teach the lessons of *Molecular Architecture*.

TABLE 14.1: 60 Minute Class Periods per Lesson

Lesson	Number of Class Periods
14.1 *Types of Bonds that Form Between Atoms*	1.0
14.2 *The Covalent Molecules of Family 2A-8A*	1.0
14.3 *Resonance*	1.0
14.4 *Electronic and Molecular Geometry*	2.0
14.5 *Molecular Polarity*	1.0

Managing Materials

The following items are needed to teach the strategies and activities described in the Teachers Edition of the FlexBook for *Molecular Architecture*.

TABLE 14.2: Molecular Architecture Materials List

Lesson	Strategy or Activity	Materials Needed
14.1	Exploration Activity	Index cards
14.2		
14.3		
14.4	Exploration Activities	4 balloons, pin, Gumdrops, toothpicks
14.5		

Multimedia Resources

You may find these additional internet resources helpful when teaching *Molecular Architecture*:

- Lesson on Chemical Bonding http://www.visionlearning.com/library/module_viewer.php?mid=55
- Polar Bears and Penguins Bonding Activity http://www.keypress.com/Documents/chemistry/SampleLessons/SmellsTG.pdf
- Examples of polar and non-polar molecules http://www.ausetute.com.au/molpolar.html
- Tutorial on Drawing Resonance Structures http://www.chem.ucla.edu/ harding/tutorials/resonance/draw_res_str.html

Possible Misconceptions

Making the FlexBook Flexible

An important advantage of the FlexBook is the ability it gives you, the teacher, to select the chapters and lessons that you think are most important for your own classes. You should also consult the standards correlation table that follows when selecting chapters and lessons to include in the FlexBook for your classes.

TABLE 14.3: Standard Addressed by the Lessons in Molecular Architecture

Lesson	California Standards	NSES Standards	AAAS Benchmarks
14.1	2a		
14.2			
14.3			
14.4	2f		
14.5	2f		

14.1 Types of Bonds that Form Between Atoms

Key Concepts

In this lesson students explore the relationship between electronegativity and bond type.

Lesson Objectives

- Given binary formulas and an electronegativity chart, students will identify the most likely bonding type (ionic, covalent, or polar covalent) for each compound.
- The students will describe a polar covalent bond and explain why it forms.

Lesson Vocabulary

bonding electron pair An electron pair found in the space between two molecules.

electronegativity The tendency of an atom in a molecule to attract shared electrons to itself.

octet rule The observation that atoms of non-metals tend to form the most stable molecules when they are surrounded by eight electrons (to fill their valence orbitals.)

polar covalent bond A covalent bond in which the electrons are not shared equally because one atom attracts them more strongly that the other.

Strategies to Engage

- Students may recall that there are two types of compounds - ionic and molecular (covalent). Review the properties of these compounds.

Strategies to Explore

- Students can think of electronegativity as the "greediness" of an atom in a molecule. Some atoms are more "greedy" for the electrons in a bond, and tend to have higher electronegativity values.
- Remind students that generally, electronegativity increases from bottom to top up a group, and from left to right across a period on the periodic table.

- Play a game with students. Cut 3×5 index cards in half. On the back of separate cards write the names and atomic numbers of the first 17 representative elements. On the front, write the electronegativity value of each element. Ask students to see if they can arrange the elements according to their positions in the periodic table using the electonegativity values only. Have them turn the cards over to see if their arrangements were correct.

Strategies to Extend and Evaluate

- Have students write a paragraph explaining the illustrations under the section "The Partial Ionic Character of Covalent Bonds" in their own words. Instruct students to correctly use each vocabulary term at least once in their paragraph.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 14.1 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Explain the differences among a covalent bond, a polar covalent bond, and an ionic bond.
2. Explain why a pair of atoms form a covalent bond rather than an ionic bond.
3. Predict which of the following bonds will be more polar and explain why; $P-Cl$ or $S-Cl$.

14.2 The Covalent Molecules of Family 2A-8A

Key Concepts

In this lesson students explore hybridization in various molecules.

Lesson Objectives

- Given binary formulas and an electronegativity chart, students will identify the most likely bonding type (ionic, covalent, or polar covalent) for each compound.
- The students will draw Lewis structures for simple molecules that violate the octet rule.
- Given a list of binary compounds, the students will identify those that require electron promotion in the explanation of their bonding.
- The students will identify the type of hybridization in various molecules.
- The students will explain the necessity for the concept of hybridized orbitals.

Lesson Vocabulary

hybrid orbitals A set of orbitals adopted by an atom in molecule, different from those of the atom in the free state.

hybridization A mixing of the native orbitals on a given atom to form special atomic orbitals for bonding.

VSEPR model A model whose main postulate is that the structure around a given atom in a molecule is determined by minimizing electron-pair repulsion.

Strategies to Engage

- A hybrid is a combination of two or more different things. Before beginning this lesson, facilitate a discussion with students about different types of hybrids they may be familiar with in areas such as mythology, biology, music, computers, transportation, and even video games.

Strategies to Explore

- This lesson includes descriptions of covalent bonding that occurs in groups 3A-8A. Before reading, prepare less proficient readers by having students write the following on the top of separate sheets of notebook paper. Instruct students to write notes as they read each section. DI **Less Proficient Readers**

The Covalent Bonds of Family 3A

The Covalent Bonds of Family 4A

The Covalent Bonds of Family 5A

The Covalent Bonds of Family 6A

The Covalent Bonds of Family 7A

The Covalent Bonds of Family 8A

Strategies to Extend and Evaluate

- Read each statement in the lesson summary. Have students indicate whether or not they understand each statement by using thumbs up or thumbs down to show "Yes" or "No". Whenever a student uses a thumbs down to show "No", use this as an opportunity to review this concept with the class. DI **English Language Learners**

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 14.2 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. What is the designation for the hybrid orbitals formed from each of the following combinations of atomic orbitals and what is the bond angle associated with the hybrid orbitals?

TABLE 14.4:

Orbitals Combined	Type of Hybridization	Bond Angles
one *s* and one *p*		
one *s* and two *p*		
one *s* and three *p*		

2. What laboratory evidence necessitates the theory of hybridized orbitals?
3. Fill in the type of hybridization necessary for the following molecules.

TABLE 14.5:

Molecule	Type of Hybridization
H_2O	
NH_3	

TABLE 14.5: (continued)

Molecule	Type of Hybridization
$BeCl_2$	
NaH	
BF_3	
PCl_5	
BrF_5	
SF_6	
XeF_2	

14.3 Resonance

Key Concepts

In this lesson students explore resonance structures of covalent molecules.

Lesson Objectives

- The student wills describe (chemistry) resonance.
- The student wills explain the equivalent bond strengths in a resonance situation.

Lesson Vocabulary

bond energy The energy required to break a given chemical bond.

bond length The distance between the nuclei of the two atoms connected by a bond.

resonance A condition occurring when more than one valid Lewis structure can be written for a particular molecule. The actual electronic structure is not represented by any one of the Lewis structures, but by the average of all of them.

Strategies to Engage

- Have a volunteer draw the Lewis structure of ozone, O_3, on the board. Draw another Lewis structure of ozone next to it. Ask students to use a show of hands to indicate which is the correct Lewis structure. Draw brackets around each structure and a double arrow in between them. Point out to the students that the two structures are equivalent and that they are called resonance structures. Explain to students that in this lesson, they will explore resonance structures.

Strategies to Explore

- Emphasize for students that the way the term resonance used in chemistry has nothing to do with the way it is used in other disciplines, and in everyday use.

Strategies to Extend and Evaluate

- On the board or chart paper, have students write a class summary of this lesson. Have one student come up with the first sentence and have student volunteers contribute sentences until the entire section has been summarized.
- Have students work in small groups to create a poster explaining the three resonance structures of carbon dioxide. Instruct students to include the three vocabulary terms on their posters.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 14.3 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Write a Lewis structure for CO_3^{2-}. Explain why all three carbon-oxygen bonds have the same length.

14.4 Electronic and Molecular Geometry

Key Concepts

In this lesson students explore the use of VSEPR theory in determining the molecular geometry of covalent compounds.

Lesson Objectives

- The students will state the main postulate in VSEPR theory.
- The students will identify both the electronic and the molecular geometry for simple binary compounds.

Lesson Vocabulary

unshared electron pair An unshared electron pair, also known as a non-bonding pair of electrons or as a lone pair of electrons, is two electrons in the same orbital in the outer shell of an atom that are not used in the formation of a covalent bond.

electronic geometry The geometric arrangement of orbitals containing the shared and unshared electron pairs surrounding the central atom of a molecule or polyatomic ion.

molecular geometry The specific three-dimensional arrangement of atoms in molecules.

Strategies to Engage

- Draw two models of water, H_2O, on the board. Draw one with a bent shape and the other with a linear shape. Tell students that a water molecule has a bent, rather than linear shape. Explain to students that in this lesson they will learn how to determine the shapes of molecules.

Strategies to Explore

- Have students build molecules using gumdrops to represent atoms and toothpicks to represent the bonds between them.
- Tie four balloons together. Use a pin to pop one balloon at a time to show how the shape changes from tetrahedral, to trigonal planar, to linear. Have students guess what the next shape will be each time a balloon is popped.

Strategies to Extend and Evaluate

- Have students write a short lesson comparing and contrasting electronic and molecular structure. They should include specific examples and illustrations in the lesson.
- Have students complete the lab named ***Molecular Models and Shapes Lab***. This lab is located in the Supplemental Lab Book.

Lesson Worksheets

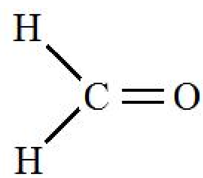
There are no worksheets for this lesson.

Review Questions

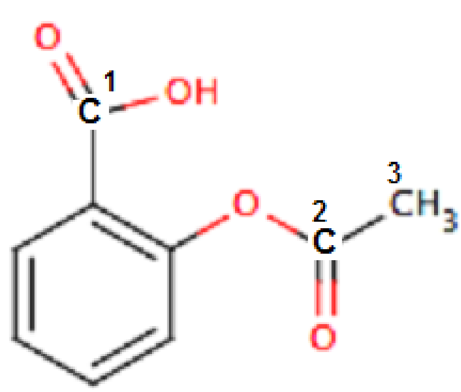
Have students answer the Lesson 14.4 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Write a Lewis structure for OF_2 that obeys the octet rule.
2. Write a Lewis structure for H_2CO that obeys the octet rule. (C is the central atom.)
3. What is the geometrical shape of the H_2CO molecule?
4. What is the bond angle inSCl_2?
5. What is the molecular shape of ICl_3?
6. What is the molecular shape of $XeCl_4$?
7. The ion I_3^- molecule has been produced in the lab but the molecule F_3^- has not. Offer an explanation as to why F_3^- cannot be produced in the lab.
8. 

The molecule shown here is methanal. What is the hybridization of the carbon atom in this molecule?

9. 

1. What is the hybridization of carbon 1?
2. What is the hybridization of carbon 2?
3. What is the hybridization of carbon 3?
4. What is the total number of pi bonds in the molecule?

14.5 Molecular Polarity

Key Concepts

In this lesson students explore how to determine molecular polarity.

Lesson Objectives

- The students will determine whether bonds are polar or non-polar.
- The students will determine whether simple molecules are polar or non-polar.

Lesson Vocabulary

polar bond A covalent bond in which the shared pair of electrons are not shared equally, owing to a difference in the electronegativity of the two atoms.

molecular symmetry The property of a molecule that enables it to undergo inversion through a line, a point, or a plane, and its new state is indistinguishable from its original state.

dipole A pair of equal and opposite charges separated by a small distance. A molecular dipole is produced when the centers of positive and negative charge do not coincide.

Strategies to Engage

- Have students read the review questions at the end of this section. This way, students will be familiar with the types of information that they will explore in this lesson.

Strategies to Explore

- Point out to students that if all of the bonds in a molecule are nonpolar, then the molecule itself is nonpolar. If the molecule has at least one polar bond, its polarity is determined by its shape.

Strategies to Extend and Evaluate

- Stand with your arms raised straight at your side so that you model a "t". Ask: If two people were pulling your hands with the same amount of strength, would you move? (No). Move your arms forward slightly. Ask: Now, if two people were pulling your hands with the same amount of strength, would you move? (Yes). Ask: In what direction would you move? (Forward). Ask students to write a paragraph relating this demonstration to the concept of molecular polarity. Ask students to use each vocabulary term at least one time in their paragraph.
- Have students organize the information explored in this lesson into a flowchart that can be used to determine if a molecule is polar or nonpolar.

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled ***Molecular Geometry.*** Ask students to complete the worksheet alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 14.5 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. What two requirements should be satisfied for a molecule to be polar?
2. BF_3 contains polar bonds but the molecule is not polar. Why not?
3. Which of the following molecules will be polar: ICl_5, $XeCl_4$, $SeCl_6$?
4. Which of the following molecules will be polar: PCl_3, SCl_2, SiF_4?
5. Which of the following molecules will have the largest dipole moment?
   a. C_2H_2
   b. CH_2Cl_2
   c. BF_3
   d. CH_3CH_2OH
   e. HF
6. Which of the following molecules will have a triple bond?
   a. C_2H_2
   b. CH_2Cl_2
   c. BF_3
   d. CH_3CH_2OH
   e. HF
7. Which of the following molecules has a central atom with sp^2 hydridized orbitals?
   a. C_2H_2
   b. CH_2Cl_2
   c. BF_3
   d. CH_3CH_2OH
   e. HF

CHAPTER 15 The Mathematics of Compounds TE

Chapter Outline

The Mathematics of Compounds

Outline

The chapter *The Mathematics of Compounds* consists of four lessons that develop the skills involved in formula stoichiometry.

- **Lesson 15.1 Determining Formula and Molecular Masses**
- **Lesson 15.2 The Mole**
- **Lesson 15.3 Percent Composition**
- **Lesson 15.4 Empirical and Molecular Formulas**

Overview

In these lessons, students will explore:

- Formula and molecular masses of compounds.
- Calculations involving the mole.
- The calculation of percent compositions given either the masses of the elements in a compound, or the chemical formula of a compound.
- Empirical and molecular formula calculations.

Science Background Information

This material is provided for teachers who are just beginning to instruct in this subject area.

- ***Avogadro's Number***

1811 was the year that Lorenzo Romano Amedeo Carlo Avogadro de Quaregna e di Cerreto - better known as Amedeo Avogadro, published his now famous hypothesis. it stated that equal volumes of gases contain the same number of particles. The nature of those particles was still a topic of considerable debate. Avogadro produced his theory based on the results of Joseph-Louis Gay-Lussac, who showed that when different gases combine, they do so in simple whole number ratios. His contemporary, John Dalton, responsible for the similar sounding Law of Multiple Proportions, reacted critically to Gay-Lussac's work. Dalton suggested that the atoms in gases were not capable of attaching; he argued that they would repel each other.Avogadro recognized that the viewpoints of both Dalton and Gay-Lussac could both be operable if, in his words, the same volume of gas contained the same number of molecules. It must be understood that the distinction between atoms and molecules did not exist in 1811 and the two words were used interchangeably.

Avogadro's principle did not gain adherents until the concept of the atom became more solidly established. Likewise, the actual determination of what has become known as Avogadro's number, was not accomplished until after Avogadro's death in 1856. Johann Josef Loschmidt, an Austrian chemist, developed a method for the first estimate of the actual number. His technique entailed measuring the difference in volume between a given liquid substance, and the volume of that material upon evaporation into the gas phase. He reasoned that in the liquid phase, all of the liquid molecules touched their adjacent molecules and that there was no empty space. Thus the total volume of the liquid was equivalent to the volume of all of the liquid molecules added together. Comparing the volumes of the liquid and gas phases, he estimated that there were about 5×10^{22} molecules in a volume of gas. By defining the number of molecules in a cubic meter of gas at standard temperature and pressure, he derived what is now known as "Loschmidt's number" or 2.686×10^{25}.

The establishment of a more carefully calculated value for the number of particles in one mole of any substance was made by Albert Einstein in the early twentieth century. Rather than using a gas model, Einstein's method was based on evaluating the number of sugar molecules in a sample of sugar water. His calculation was based on the average velocity that the individual molecules diffused through a membrane and was initially published as 2.1×10^{23}. Several years later, with new data from more accurate measurements, he redefined the value as 6.1×10^{23}.

The value now used in chemistry texts, 6.022×10^{23}, was arrived at by a different technique. The current value has been calculated by using x-ray diffraction in crystal lattices of silicon atoms.

Pacing the Lessons

Use the table below as a guide for the time required to teach the lessons of *The Mathematics of Compounds.*

TABLE 15.1: 60 Minute Class Periods per Lesson

Lesson	Number of Class Periods
15.1 *Determining Formula and Molecular Masses*	2.0
15.2 *The Mole*	2.0
15.3 *Percent Composition*	1.5
15.4 *Empirical and Molecular Formulas*	2.0

Managing Materials

The following items are needed to teach the strategies and activities described in the Teachers Edition of the FlexBook for *The Mathematics of Compounds.*

TABLE 15.2: The Mathematics of Compounds Materials List

Lesson	Strategy or Activity	Materials Needed
15.1		
15.2		
15.3		
15.4		

Multimedia Resources

You may find these additional web based resources helpful when teaching *The Mathematics of Compounds.*

- "Mole Day" activities: http://www.moleday.org/
- Interactive mole quiz: http://glencoe.com/qe/science.php?qi=978
- Mole Conversion practice problem generator: http://science.widener.edu/svb/tutorial/massmoles.html
- "Chemical Composition" flashcards: http://college.cengage.com:80/chemistry/general/zumdahl/world_of_chem/1e/students/flashcards/ch06/index.html

Possible Misconceptions

Making the FlexBook Flexible

An important advantage of the FlexBook is the ability it gives you, the teacher, to select the chapters and lessons that you think are most important for your own classes. You should also consult the standards correlation table that follows when selecting chapters and lessons to include in the FlexBook for your classes.

TABLE 15.3: Standard Addressed by the Lessons in The Mathematics of Compounds

Lesson	California Standards	NSES Standards	AAAS Benchmarks
15.1			
15.2	$3b, 3c, 3d$		
15.3	$3d$		
15.4			

15.1 Determining Formula and Molecular Mass

Key Concepts

In this lesson students explore formula and molecular masses of compounds.

Lesson Objectives

- When given the formula or name of a compound and a periodic table, the student will be able to calculate the formula mass.

Lesson Vocabulary

formula mass The sum of the atomic masses of the atoms in a formula.

molecular mass The mass of a molecule found by adding the atomic masses of the atoms comprising the molecule.

Strategies to Engage

- Review mass number and atomic number with students. Remind students that nearly the entire mass of an atom is determined by the protons and neutrons, and that the mass number of an atom is the sum of its protons and neutrons. Review with students how the atomic mass of an element is calculated by a weighted average of the atoms in a naturally occurring sample of the element.
- Write the atomic mass of ten different elements on the board. Time students as they locate the elements on the periodic table and write down the name, chemical symbol, and atomic number for each one. Award a prize to the student who completes the list in the least amount of time.

Strategies to Explore

- Have students create a Venn diagram of formula mass vs. molecular mass.
- Use the mathematical calculations in this lesson to reduce the reliance on language skills. As you go through each example problem, use them to explain the concepts explored in this chapter. DI **English Language Learners**

Strategies to Extend and Evaluate

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Calculating Molar Mass Worksheet**. Ask students to complete the worksheet alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 15.4 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Calculate the formula mass for each of the following.
   a. K_2SO_4
   b. CuO
   c. $Mg_3(AsO_4)_2$
   d. $Ca_3(PO_4)_2$
   e. Fe_2O_3
   f. $Al(OH)_3$
   g. $(NH_4)_2S$
   h. $C_{12}H_{22}O_{11}$
2. How many times heavier are bromine atoms on the average than neon atoms?
3. An unknown element, M, combines with oxygen to form a compound with a formula of MO_2. If 25.0 grams of the unknown element combines with 4.50 grams of oxygen, what is the atomic mass of M?

15.2 The Mole

Key Concepts

In this lesson students explore calculations involving the mole.

Lesson Objectives

- Given the number of particles of a substance, the student will use Avogadro's number to convert to moles and vice versa.
- Given the number of moles of a substance, the student will use the molar mass to convert to grams and vice versa.

Lesson Vocabulary

Avogadro's number The number of objects in a mole; equal to 6.02×10^{23}.

mole An Avogadro's number of objects.

Strategies to Engage

- Explain to students that in this lesson, they will be introduced to the mole, which is the SI unit that describes the amount of a substance. Write Avogadro's number in standard form on the board. If students comment that the mole is a very large quantity, remind them that atoms are incredibly tiny. Tell students that just as a dozen eggs is 12 eggs, a mole of eggs is 6.02×10^{23} eggs.

Strategies to Explore

- Use the mathematical calculations in this lesson to reduce the reliance on language skills. As you go through each example problem, use them to explain the concepts explored in this lesson. DI **English Language Learners**

Strategies to Extend and Evaluate

- Encourage interested students to research the work of Amedeo Avogadro and his contributions to chemistry. Students should be prepared to share their findings with the class.
- Have students create a "mole hill" and display it on a wall of the classroom. Have each student contribute two or three quantities that are equal to a mole such as 32.1 *g* of sulfur, 18.0 *g* of water, or 34.0 *g* of hydrogen peroxide. Collect the students' contributions and use them to create a "mole hill". You may want to ask a student to draw a large mole (the animal) to sit on top of the hill.

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Moles Worksheet**. Ask students to complete the worksheet alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 15.2 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Convert the following to moles.
   a. 60.0 grams of NaOH
   b. 2.73 grams of NH_4Cl
   c. 5.70 grams of ZrF_4
   d. 10.0 grams ofPbO_2
2. Convert the following to grams.
   a. 0.100 moles of CO_2
   b. 0.437 moles of NaOH
   c. 0.500 moles of $(NH_4)_2CO_3$
   d. 3.00 moles of $ZnCl_2$
3. How many molecules are present in the following quantities?
   a. 0.250 mole of H_2O
   b. 6.00 moles of H_2SO_4
   c. 0.00450 mole of $Al_2(CO_3)_3$
4. How many moles are present in the following quantities?
   a. 1.00×10^{20} molecules of H_2O
   b. 1.00×10^{25} molecules of H_2
   c. $5,000,000,000,000$ atoms of carbon
5. How many molecules (or atoms) are present in the following masses?
   a. 1.00 gram of Na_2CO_3
   b. 8.00 grams of helium
   c. 1000. grams of H_2O

6. Convert the following to grams.
   a. 1.00×10^{23} molecules of H_2
   b. 1.00×10^{24} molecules of $AlPO_4$
   c. 1.00×10^{22} molecules of NaOH
7. What is the mass of a single atom of silver, Ag, in grams?
8. If you had a silver bar that exactly fit inside a 100. mL graduated cylinder and filled it precisely to the 100. mL mark, how many atoms of silver would be in the bar? The density of silver is 10.5 g/mL.

15.3 Percent Composition

Key Concepts

In this lesson students explore the calculations of percent composition given either the masses of the elements in a compound or the chemical formula of a compound.

Lesson Objectives

- Given either masses of elements in a compound, the student will calculate the percent composition by mass.
- Given the formula or name of a compound, the student will calculate the percent composition by mass.

Lesson Vocabulary

percent composition The proportion of an element present in a compound found by dividing the mass of the element by the mass of the whole compound and multiplying by 100.

Strategies to Engage

- Ask "How could I determine the percentage of males and females in the classroom?"(Count the number of males and the number of females. Add them up, then divide the number of males by the total, and the number of females by the total. Then, multiply each by 100). Explain to students that in this lesson they will explore percent composition, which is the percent by mass of each element in a compound. It is calculated in much the same way.

Strategies to Explore

- Use the mathematical calculations in this lesson to reduce the reliance on language skills. As you go through each example problem, use them to explain the concepts explored in this lesson. DI **English Language Learners**

Strategies to Extend and Evaluate

- Ask small groups of students to write a four-step method of calculating percent composition. Choose the best set of steps out of all of the groups. Write the steps on the board, and use those steps to complete practice and

review problems.
- Write the formulas for different compounds on strips of paper and place them in a hat or container. Have students draw papers from the hat and then calculate the percent composition of the compound they have drawn.

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Percent Composition Worksheet**. Ask students to complete the worksheet alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 15.3 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Determine the percent composition of the following compounds.
2. BF_3
3. $Ca(C_2H_3O_2)_2$
4. FeF_3
5. $CrCl_3$
6. $(NH_4)_3PO_4$

15.4 Empirical and Molecular Formulas

Key Concepts

In this lesson students explore empirical and molecular formula calculations.

Lesson Objectives

- The student will reduce molecular formulas to empirical formulas.
- Given either masses or percent composition of a compound, the student will determine the empirical formula.
- Given either masses or percent composition of a compound and the molar mass, the student will determine the molecular formula.

Lesson Vocabulary

empirical formula The formula giving the simplest ratio between the atoms of the elements present in a compound.

molecular formula A formula indicating the actual number of each kind of atom contained in a molecule.

Strategies to Engage

- Have students write down the lesson objectives, leaving about 5 or 6 lines of space in between. As you explore the lesson, have students write specific examples of each objective.

Strategies to Explore

- Have students create a Venn diagram of empirical formula vs. molecular formula.
- Use the mathematical calculations in this lesson to reduce the reliance on language skills. As you go through each example problem, use them to explain the concepts explored in this lesson. DI **English Language Learners**

Strategies to Extend and Evaluate

- Have students create study cards of the calculations explored in this chapter.

- Have students complete the labs ***Empirical Formula of Magnesium Oxide*** and **'Formula of a Hydrate.** These labs are located in the Supplemental Lab Book.

Lesson Worksheets

Copy and distribute the lesson worksheets in the *CK-12 Chemistry Workbook* titled **Empirical Formulas** and **Molecular Formulas**. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 15.4 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. What is the empirical formula for C_8H_{18}?
2. What is the empirical formula for C_6H_6?
3. What is the empirical formula for WO_2?
4. A compound has the empirical formula C_2H_8N and a molar mass of 46 grams/mol. What is the molecular formula of this compound?
5. A compound has the empirical formula C_2H_4NO. If its molar mass is 116.1 grams/mole, what is the molecular formula of the compound?
6. A sample of pure indium chloride with a mass of 0.5000 grams is found to contain 0.2404 grams of chlorine. What is the empirical formula of this compound?
7. Determine the empirical formula of a compound that contains 63.0 grams of rubidium and 5.90 grams of oxygen.
8. Determine the empirical formula of a compound that contains 58.0%Rb, 9.50%N, and 32.5%O.
9. Determine the empirical formula of a compound that contains 33.3%Ca, 40.0%O, and 26.7%S.
10. Find the molecular formula of a compound with percent composition 26.7%P, 12.1%N, and 61.2%Cl and with a molecular mass of 695 g/mol.

Chapter 16 Reactions TE

Chapter Outline

Unit 5 Reactions and Stoichiometry

Outline

This unit, *Reactions and Stoichiometry*, includes two chapters that introduce students to chemical reactions, chemical equations, and stoichiometric relationships.

- **Chapter 16 Chemical Reactions**
- **Chapter 17 Mathematics and Chemical Equations**

Overview

Chemical Reactions

This chapter develops the skills involved in mass and molecule to mole calculations and the determination of reaction types.

Mathematics and Chemical Equations

This chapter develops the skills involved in equation stoichiometry including limiting reactant equations, yields, and introduces heat of reaction.

Chemical Reactions

The chapter *Chemical Reactions* consists of three lessons that develop the skills involved in mass and molecule, to mole calculations and the determination of reaction types.

- **Lesson 16.1 Chemical Equations**
- **Lesson 16.2 Balancing Equations**
- **Lesson 16.3 Types of Reactions**

Overview

In these lessons, students will explore:

- The symbol equations and word equations used to describe chemical reactions.
- The balancing of chemical equations.
- Different types of chemical reactions.

Science Background Information

Fireworks

If you enjoyed a Fourth of July evening pyrotechnic display, or perhaps witnessed a New Years' Eve event, you've witnessed the results of over a thousand years' worth of research and development into the art of fireworks. The first efforts were produced in China initially by accident as they observed that when saltpeter (potassium nitrate, KNO_3) was dropped into a charcoal fire, the mixture "popped" and produced an interesting flame color. Later, as a means to surprise their enemies in battle, the earliest "Shock and Awe" campaigns featured a mixture of saltpeter, charcoal and sulfur. The mechanism by which fireworks operate involves heating the proper ratio of these materials (75% KNO_3, 15% carbon and 10% sulfur), and generating a chemical reaction to produce nitrogen and carbon dioxide gases. These initial "gung pow" were mainly explosive devices directed into the air, but later new additions brought whistling sound effects and a spectrum of colors to dazzle their opponents. The energy needed to propel the shell and to excite the composite atoms is still provided by a gunpowder formula.

FIGURE 16.1

The brilliant colors that produce the *oohs* and *aahs* of today's displays are mainly due to elements like magnesium, which results in a blinding white effect. On an atomic level, the energy imparted by the explosion causes the atom's electrons to be promoted to a higher energy level. When the atoms relax back to the ground state, a specific amount of energy is released and the color of visible light reveals the frequency of light corresponding to that energy value. The red coloration is due to the presence of lithium or strontium salts such as lithium or strontium carbonate. Sodium salts (usually nitrate) generate a yellow hue and calcium chloride or sulfate result in orange coloration.

Barium chloride supplies a green color. The all-American red, white and blue display is difficult to construct due to the complexity of finding a blue colored explosive. Usually copper chloride in a blue-violet hue is substituted. This copper salt's instability at the high temperatures of the exploding device has caused modern day pyrotechnical researchers to continue the search for a reliable source of blue color.

The shape of the image produced when the shell explodes in the air is a function of how the components are arranged in the shell. When the pyrotechnic device explodes as the resultant gases are produced, the arrangement of the salts in the mortar shell will mirror the pattern produced by the explosion. Dividing the materials into different compartments can also produce a "time – delay" effect, where the display effects occur sequentially.

The sound effects employed as a counterpoint to the visual display are also the result of chemical reactions. Adding bismuth trioxide to the mixture generates "popping" noises, whereas copper salicylate yields a "whistling" sound.

On the next occasion when a fireworks display rises to the sky, you'll not only enjoy the beautiful visual effects, but have an appreciation for the science that went into the presentation.

Pacing the Lessons

Use the table below as a guide for the time required to teach the lessons of *Chemical Reactions.*

TABLE 16.1: 60 Minute Class Periods per Lesson

Lesson	Number of Class Periods
16.1 *Chemical Equations*	0.5
16.2 *Balancing Equations*	2.0
16.3 *Types of Reactions*	2.0

Managing Materials

The following items are needed to teach the strategies and activities described in the Teachers Edition of the FlexBook for *Chemical Reactions.*

TABLE 16.2: Chemical Reactions Materials List

Lesson	Strategy or Activity	Materials Needed
16.1		
16.2		
16.3	Extension Activity	Index cards

Multimedia Resources

You may find these additional web based resources helpful when teaching *Chemical Reactions*:

- Balancing chemical equations activity: http://www.middleschoolscience.com/balance.html
- Balancing chemical equations game: http://funbasedlearning.com/chemistry/chemBalancer/default.htm

Possible Misconceptions

Making the FlexBook Flexible

An important advantage of the FlexBook is the ability it gives you, the teacher, to select the chapters and lessons that you think are most important for your own classes. You should also consult the standards correlation table that follows when selecting chapters and lessons to include in the FlexBook for your classes.

TABLE 16.3: Standard Addressed by the Lessons in Chemical Reactions

Lesson	California Standards	NSES Standards	AAAS Benchmarks
16.1	3a		
16.2	3a		
16.3			

16.1 Chemical Equations

Key Concepts

In this lesson, students will explore the word equations and symbol equations that chemists use to describe chemical reactions.

Lesson Objectives

- The student will read chemical equations, and provide requested information contained in the equation including information about substances, reactants, products, and physical states.
- The student will convert symbolic equations into word equations and vice versa.
- The student will use the common symbols, $+, (s), (L), (g), (aq)$, and $\rightarrow$ appropriately.
- The student will describe the roles of subscripts and coefficients in chemical equations.
- The student will balance chemical equations with the simplest whole number coefficients.

Lesson Vocabulary

reactants The substances on the left side of a chemical equation.

products The substances on the right side of a chemical equation.

As a subscript to a formula, indicates the substance is in the solid phase.

As a subscript to a formula, indicates the substance is in the liquid phase.

As a subscript to a formula, indicates the substance is in the gaseous phase.

As a subscript to a formula, indicates the substance is dissolved in water.

Strategies to Engage

- Review with students the difference between physical and chemical changes. Remind students that in a chemical change (reaction) new substances are formed. These new substances have different properties than the original substance.

Strategies to Explore

- Perform the ***Briggs-Rauscher Oscillating Reaction*** *demonstration. This demonstration is located in the Supplemental Lab Book.*

Strategies to Extend and Evaluate

- Have each student share an example of a chemical reaction that they see in everyday life. Have each student identify the products, reactants, and physical states of all substances involved and discuss these with the class.
- Have students write the equation for the reaction of solid sodium bicarbonate with hydrochloric acid to produce aqueous sodium chloride, water, and carbon dioxide gas. Have them label each vocabulary word in the appropriate place in the equation.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 16.1 Review Questions that are listed at the end of the lesson in their Flexbook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Mothballs are commonly used to preserve clothing in "off-season." We recognize mothballs due to its smell because of a chemical compound known as Naphthalene, $C_{10}H_8$. What are the different elements found in naphthalene and how many atoms of each are found in the formula?
2. Do you think a chemical reaction occurs every time two substances are placed together in a reaction vessel?
3. Transfer the following symbolic equations into word equations.
   a. $H_2SO_{4(aq)} + NaCN_{(aq)} \rightarrow HCN_{(aq)} + Na_2SO_{4(aq)}$
   b. $Cu(s) + AgNO_{3(aq)} \rightarrow Ag_{(s)} + Cu(NO_3)_{2(aq)}$
   c. $Fe_{(s)} + O_{2(g)} \rightarrow Fe_2O_{3(s)}$
4. Transfer the following equations from word equations into symbolic equations.
   a. Solid calcium metal is placed in liquid water to produce aqueous calcium hydroxide and hydrogen gas.
   b. Gaseous sodium hydroxide is mixed with gaseous chlorine to produce aqueous solutions of sodium chloride and sodium hypochlorite plus liquid water.
   c. Solid xenon hexafluoride is mixed with liquid water to produce solid xenon trioxide and gaseous hydrogen fluoride.
5. Did you know that you can simulate a volcanic eruption in a lab that looks like the real thing? A source of heat is gently placed it into a mound of ammonium dichromate. The ammonium dichromate decomposes to solid chromium (III) oxide, nitrogen monoxide gas, and water vapor. Write the symbolic reaction for the "volcanic eruption".

16.2 Balancing Equations

Key Concepts

In this lesson students will learn to balance non-redox chemical equations.

Lesson Objectives

- Demonstrate the Law of Conservation of Matter in a chemical equation.
- Explain the roles of coefficients and subscripts in a chemical equation.
- Balance equations using the simplest whole number coefficients.

Lesson Vocabulary

law of conservation of matter Matter is neither created nor destroyed in chemical reactions.

skeletal equation A chemical equation before it has been balanced.

balanced chemical equation A chemical equation in which the number of each type of atom is equal on the two sides of the equation.

Strategies to Engage

- Review with students the Law of Conservation of Matter. Explain to students that in this lesson they will demonstrate the relationship of this law to chemical equations.

Strategies to Explore

- Have students create a Venn diagram comparing and contrasting coefficients and subscripts.

Strategies to Extend and Evaluate

- Divide students into groups of four. Within each group of four, have pairs of students take turns coming up with equations for the other pair of students to balance.

- Have interested students develop a way to teach younger kids how to balance equations using candies or other small objects. Students should be prepared to demonstrate their method in front of the class using real examples.

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Balancing Equations**. Ask students to complete the worksheet alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 16.2 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Explain in your own words why it is essential that subscripts remain constant but coefficients can change.
2. Which set of coefficients will properly balance the following equation?

$$C_2H_6 + O_2 \rightarrow CO_2 + H_2O$$

   a. $1,1,1,1$
   b. $1,3,2,2$
   c. $1,3.5,2,3$
   d. $2,7,4,6$

3. When properly balanced, what is the sum of all the coefficients in the following chemical equation?

$$SF_4 + H_2O \rightarrow H_2SO_3 + HF$$

   a. 4
   b. 7
   c. 9
   d. None of the above

4. When the following equation is balanced, what is the coefficient found in front of the O_2?

$$P_4 + O_2 + H_2O \rightarrow H_3PO_4$$

   a. 1
   b. 3
   c. 5
   d. 7

5. Balance the following equations.
   a. $XeF_{6(s)} + H_2O_{(L)} \rightarrow XeO_{3(s)} + HF_{(g)}$
   b. $Cu_{(s)} + AgNO_{3(aq)} \rightarrow Ag_{(s)} + Cu(NO_3)_{2(aq)}$
   c. $Fe_{(s)} + O_{2(g)} \rightarrow Fe_2O_{3(s)}$
   d. $Al(OH)_3 + Mg_3(PO_4)_2 \rightarrow AlPO_4 + Mg(OH)_2$

16.3 Types of Reactions

Key Concepts

In this lesson, students will learn to identify the different types of chemical reactions.

Lesson Objectives

- Identify the types of reactions.
- Predict the products in different types of reactions.
- Distinguish between the different types of reactions.
- Write balanced chemical equations and identify the reaction type given only the reactants.

Lesson Vocabulary

synthesis A synthesis reaction is one in which two or more reactants combine to make one type of product. $(A + B \rightarrow C)$.

decomposition A decomposition reaction is one in which one type of reactant breaks down to form two or more products. $(C \rightarrow A + B)$.

single replacement (metal) In a single replacement (metal) reaction, one element replaces the metal cation of the compound reactant to form products. $(A + BC \rightarrow AC + B)$.

single replacement (many metals with acid) In a single replacement (many metals with acid) reaction, one element replaces the hydrogen cation of the compound (which is an acid) reactant to form products. Example: $A + 2\ HC \rightarrow AC_2 + H_2$.

single replacement (non-metal) In a single replacement (non-metal) reaction, one element replaces the non-metal (anion) of the compound reactant to form products. $(XY + Z \rightarrow XZ + Y)$.

double replacement For double replacement reactions two reactants will react by having the cations replace the anions. $(AB + XY \rightarrow AY + XB)$. Double replacement reactions are also called *metathesis* reactions sometimes.

combustion (complete) Combustion is the burning in oxygen, usually a hydrocarbon. $(\text{fuel} + O_2 \rightarrow CO_2 + H_2O)$.

combustion (incomplete) Incomplete combustion is the inefficient burning in oxygen, usually a hydrocarbon. Inefficient burning means there in not enough oxygen to burn all of the hydrocarbon present, sometimes carbon (soot) is also a side product of these reactions. $(\text{fuel} + O_2 \rightarrow CO_2 + H_2O)$.

hydrocarbons Compounds containing hydrogen and carbon.

Strategies to Engage

- Before beginning this lesson, write four or five of the sample questions found in the student book that require students to predict the products of chemical reactions. Explain to students that by the end of this lesson, they will be able to predict the products of these and other chemical reactions.

Strategies to Explore

- Perform the ***Explosive Mixtures of Ethyne and Air*** *demonstration. This demonstration is located in the Supplemental Lab Book.*

Strategies to Extend and Evaluate

- Have students write each of the five types of reactions explored in this lesson on separate index cards. Have groups of students take turns picking a card and acting out the type of reaction on the card, using props if necessary, while the others guess which one it is.
- As a review of the chapter vocabulary, suggest that students make flash cards, with the vocabulary term on one side, and a definition and example of it on the other.
- Have students make a poster of each of the types of chemical reactions explored in this chapter. Ask students to include specific examples of each reaction.
- Have students complete the labs **Chemical Reactions in Microscale, Chemical Reactions Using Probeware** *and* **Types of Chemical Reactions.** These labs are located in the Supplemental Lab Book.

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Types of Chemical Reactions**. Ask the students to complete the worksheet individually or in groups.

Review Questions

Have students answer the Lesson 16.3 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. When balancing combustion reactions, did you notice a consistency relating to whether the number of carbons in the hydrocarbon was odd or even?
2. Distinguish between synthesis and decomposition reactions.
3. When dodecane, $C_{10}H_{22}$, burns in excess oxygen, the products would be

   a. $CO_2 + 2\,H_2$
   b. $CO + H_2O$
   c. $CO_2 + H_2O$

d. CH_4O_2

4. In the decomposition of antimony trichloride, which of the following products and quantities will be found?

a. $An+Cl_2$
b. $2\,An+3\,Cl_2$
c. $Sb+Cl_2$
d. $2\,Sb+3\,Cl_2$

5. Acetylsalicylic acid (Aspirin^{TM}), $C_9H_8O_4(s)$, is produced by reacting acetic anhydride, $C_4H_6O_3(l)$, with salicylic acid, $C_7H_6O_3(s)$. The other product in the reaction is water. Write the balanced chemical equation.
6. When iron rods are placed in liquid water, a reaction occurs. Hydrogen gas evolves from the container and iron(III) oxide forms onto the iron rod.

a. Write a balanced chemical equation for the reaction.
b. What type of reaction is this?

7. A specific fertilizer is being made at an industrial plant nearby. The fertilizer is called a triple superphosphate and has a formula $Ca(H_2PO_4)_2$. It is made by sand and clay that contains phosphate and then treating it with a calcium phosphate solution and phosphoric acid. The simplified reaction is calcium phosphate reacting with the phosphoric acid to yield the superphosphate. Write the balanced chemical reaction and name the type.

16.4 References

1. Semnoz. http://commons.wikimedia.org/wiki/File:200508-DSCN0417Fireworks.jpg. CC-BY-SA 3.0

Chapter 17 Mathematics and Chemical Equations TE

Chapter Outline

Mathematics and Chemical Equations

Outline

The chapter *Mathematics and Chemical Equations* consists of five lessons that develop the skills involved in equation stoichiometry including limiting reactant equations, yields, and introduces heat of reaction.

- **Lesson 17.1 The Mole Concept and Equations**
- **Lesson 17.2 Mass-Mass Calculations**
- **Lesson 17.3 Limiting Reactant**
- **Lesson 17.4 Percent Yield**
- **Lesson 17.5 Energy Calculations**

Overview

In these lessons, students will explore:

- Mole relationships in balanced chemical equations.
- Mass relationships in balanced chemical equations.
- Limiting and excess reactants in chemical reactions.
- Theoretical, actual, and percent yield of a product.
- Energy changes in chemical processes.

Science Background Information

This information is provided for teachers who are just beginning to instruct in this subject area.

Excess and Limiting Reactants

The chemical name for chalk is calcium carbonate. The reaction between sodium carbonate and calcium chloride produces calcium carbonate.

$$Na_2CO_{3(aq)} + CaCl_{2(aq)} \rightarrow CaCO_{3(s)} + 2\,NaCl_{(aq)}$$

Stoichiometry allows us to compare the amounts of various species involved in a reaction. In order to determine which of the reactants is the limiting reactant, we must take into account both the amounts present and how they relate stoichiometrically in the balanced equation. Why do chemists use limiting reactants? The reason lies in the fact that not all reactions go to 100% completion; in fact the majority of the really interesting ones do not. However, scientists can use an equilibrium "trick" to get the stubborn reactions to go to completion. They start with an excess of one of the reactants to "push" the reaction to make more product. This essentially makes the other reactant the limiting reactant.

Pacing the Lessons

Use the table below as a guide for the time required to teach the lessons of *Mathematics and Chemical Equations*.

TABLE 17.1: 60 Minute Class Periods per Lesson

Lesson	Number of Class Periods
17.1 *The Mole Concept and Equations*	1.0
17.2 *Mass-Mass Calculations*	2.0
17.3 *Limiting Reactant*	1.5
17.4 *Percent Yield*	1.5
17.5 *Energy Calculations*	1.5

Managing Materials

The following items are needed to teach the strategies and activities described in the Teachers Edition of the FlexBook for *Mathematics and Chemical Equations*.

TABLE 17.2: Mathematics and Chemical Equations Materials List

Lesson	Strategy or Activity	Materials Needed
17.1		
17.2		
17.3	Engagement Activity	5 plates, 4 forks, 4 knives, and 3 spoons
17.4		
17.5	Engagement Activity	2 – 50 mL beakers, 2 g of $NaOH$, 2 g of $NaHCO_3$, 2 thermometers

Multimedia Resources

You may find these additional web based resources helpful when teaching *Mathematics and Chemical Equations*.

- Stoichiometry game: http://www.chemcollective.org/mr/
- Humorous "Mole" Video: http://www.youtube.com/watch?v=1R7NiIum2TI

Possible Misconceptions

Making the FlexBook Flexible

An important advantage of the FlexBook is the ability it gives you, the teacher, to select the chapters and lessons that you think are most important for your own classes. You should also consult the standards correlation table that follows when selecting chapters and lessons to include in the FlexBook for your classes.

TABLE 17.3: Standard Addressed by the Lessons in Mathematics and Chemical Equations

Lesson	California Standards	NSES Standards	AAAS Benchmarks
17.1	3c, 3e		
17.2	3e		
17.3	3e		
17.4	3f		
17.4	7b		

17.1 The Mole Concept and Equations

Key Concepts

- In this lesson, students will explore mole relationships in balanced chemical equations.

Lesson Objectives

- Express chemical equations in terms of molecules, formula units, and moles.
- Determine mole ratios in chemical equations.
- Explain the importance of balancing equations before determining mole ratios.
- Use mole ratios in balanced chemical equations.

Lesson Vocabulary

chemical coefficient The number in front of a molecule's symbol in a chemical equation indicates the number molecules participating in the reaction. If no coefficient appears, we interpret it as meaning 1.

formula unit The empirical formula of an ionic or covalent compound.

stoichiometry The calculation of quantitative relationships of the reactants and products in a balanced chemical reaction. Sometimes it is called reaction stoichiometry to distinguish it from composition stoichiometry.

Strategies to Engage

- Review balancing equations by having students write the steps to balancing equations on the board. Then have them use the steps to balance an actual equation.
- Review with students that a balanced equation has numbers called coefficients in front of the chemical formulas. If there is no coefficient, it is assumed to be 1. Explain to students that in this chapter they will use the coefficients from balanced equations to calculate the quantities of reactants or products in chemical reactions.

Strategies to Explore

- Facilitate a discussion with students about the similarities between a balanced chemical equation and a recipe. Students can think of balanced chemical equations as the recipes that chemists follow to produce products from reactants.

- Use the mathematical calculations in this lesson to reduce the reliance on language skills. As you go through each example problem, use them to explain the concepts explored in this lesson. DI **English Language Learners**

Strategies to Extend and Evaluate

- Read each statement in the lesson summary. Have students indicate whether or not they understand each statement by using thumbs up or thumbs down to show "Yes" or "No". Whenever a student uses a thumbs down to show "No", use this as an opportunity to review this concept with the class. DI **English Language Learners**

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the lesson 17.1 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Distinguish between formula unit, molecular unit, and mole. Use an example in your answer.
2. Why is balancing the equation so important before the mol ratios are to be determined?
3. Is it ever possible to use mol ratios when the chemical equation has not been balanced?
4. Given the reaction between ammonia and oxygen to produce nitrogen monoxide, how many moles of water vapor can be produced from 2 mols of ammonia?

$$4\ NH_{3(g)} + 5\ O_{2(g)} \rightarrow 4\ NO_{(g)} + 6\ H_2O_{(g)}$$

   a. 3 mol
   b. 6 mol
   c. 12 mol
   d. 24 mol
5. When properly balanced, how many moles of bismuth(III) oxide can be produced from 0.625 mol of bismuth?

$$Bi_{(s)} + O_{2(g)} \rightarrow Bi_2O_{3(s)}$$

   a. 0.313 mol
   b. 0.625 mol
   c. 1 mol
   d. 1.25 mol
   e. 2 mol
6. For the following reaction, balance the equation and then determine the mole ratio of moles of $B(OH)_3$ to moles of water?

$$B_2O_{3(s)} + H_2O_{(L)} \rightarrow B(OH)_{3(s)}$$

a. 1 : 1
b. 2 : 3
c. 3 : 2
d. None of the above

7. Write the balanced chemical equation for the reactions below. When written, find the mol ratios indicated following each reaction.
   a. Gaseous propane (C_3H_8) combusts to form gaseous carbon dioxide and water; find ratio of mol O_2 to mol CO_2.
   b. Solid lithium reacts with an aqueous solution of aluminum chloride to produce aqueous lithium chloride and solid aluminum; find ratio of mol $AlCl_{3(aq)}$ to mol $LiCl(aq)$.
   c. Gaseous ethane (C_2H_6) combusts to form gaseous carbon dioxide and water; find ratio of mol $CO_{2(g)}$ to mol $O_{2(g)}$.
   d. An aqueous solution of ammonium hydroxide reacts with an aqueous solution of phosphoric acid to produce aqueous ammonium phosphate and water; find ratio of mol $H_3PO_{4(aq)}$ to mol $H_2O(l)$.
   e. Solid rubidium reacts with solid phosphorous to produce solid rubidium phosphide; find ratio of mol Rb(s) to mol P(s).
8. For the given reaction (unbalanced),

$$Ca_3(PO_4)_2 + SiO_2 + C \rightarrow CaSiO_3 + CO + P$$

   a. how many moles of silicon dioxide are required to react with 0.35 mol of carbon?
   b. how many moles of calcium phosphate are required to produce 0.45 mol of calcium silicate?
9. For the given reaction (unbalanced),

$$FeS + O_2 \rightarrow Fe_2O_3 + SO_2$$

   a. how many moles of iron(III) oxide are produced from 1.27 mol of oxygen?
   b. how many moles of iron(II) sulfide are required to produce 3.18 mol of sulfur dioxide?
10. Write the following balanced chemical equation. Ammonia and oxygen are allowed to react in a closed container to form nitrogen and water. All species present in the reaction vessel are in the gas state.
   a. how many moles of ammonia are required to react with 4.12 mol of oxygen?
   b. how many moles of nitrogen are produced when 0.98 mol of oxygen are reacted with excess ammonia?

17.2 Mass-Mass Calculations

Key Concepts

In this lesson, students learn to calculate mass relationships in balanced chemical reactions.

Lesson Objectives

- The student will define stoichiometry.
- Given the mass of one reactant or product, the student will calculate the mass of any other reactant or product.
- The student will use the factor-label method in mass-mass calculations.

Lesson Vocabulary

stoichiometry The calculation of quantitative relationships of the reactants and products in a balanced chemical reaction. Sometimes it is called reaction stoichiometry to distinguish it form composition stoichiometry.

Strategies to Engage

- Explain to students that in the last lesson they explored mole relationships in balanced equations. In this lesson they will explore mass relationships in balanced equations. Have students read the introduction, lesson vocabulary and lesson objectives, then facilitate a discussion with students about what they think will be some similarities and differences between these two concepts.

Strategies to Explore

- Point out to students the importance of writing the correct units throughout the problem, and that in the end, all of the units must cancel except for the desired unit. This will prevent students from bypassing the mole ratio and attempting to convert the mass of the given substance directly to the mass of the desired substance.
- Divide students into groups of three or four to work on problems in this lesson. Assign one student in each group to serve as a reminder to the rest of the group members to include the proper units throughout the problems.

Strategies to Extend and Evaluate

- Have students complete the lab ***Mass-Mass Relationships in a Chemical Change***. This lab is located in the Supplemental Lab Book.

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Mass-Mass Calculations**. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 17.2 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Why is it important to have a balanced chemical equation before you begin the calculations involving mass-mass conversions?
2. The standard length of a football field is 360. feet. Use the factor label method to convert this length into centimeters.
3. Given the reaction between copper (II) sulfide and nitric acid, how many grams of nitric acid will react with 2.00 g of copper(II) sulfide?

$$3\ CuS_{(s)} + 8\ HNO_{3(aq)} \rightarrow 3\ Cu(NO_3)_{2(aq)} + 2\ NO_{(g)} + 4\ H_2O_{(L)} + 3\ S_{(s)}$$

   a. 0.49 g
   b. 1.31 g
   c. 3.52 g
   d. 16.0 g

4. When properly balanced, what mass of iodine was needed to produce 2.5 g of sodium iodide in the equation below?

$$I_{2(aq)} + Na_2S_2O_{3(aq)} \rightarrow Na_2S_4O_{6(aq)} + NaI_{(aq)}$$

   a. 1.0 g
   b. 2.1 g
   c. 2.5 g
   d. 8.5 g

5. Donna was studying the following reaction for a stoichiometry project.

$$S_{(s)} + 3F_{2(g)} \rightarrow SF_{6(s)}$$

   She wondered how much she could obtain if she used 3.5 g of fluorine. First she has to balance the equation. What mass of $SF_6(s)$ would she obtain from the calculation using this amount of fluorine?

   a. 3.5 g
   b. 4.5 g

   c. 10.5 g
   d. 13.4 g

6. Aqueous solutions of aluminum sulfate and sodium phosphate are placed in a reaction vessel and allowed to react. The products of the reaction are aqueous sodium sulfate and solid aluminum phosphate.
   a. Write a balanced chemical equation to represent the above reaction.
   b. How many grams of sodium phosphate must be added to completely react all of 5.00 g of aluminum sulfate?
   c. If 3.65 g of sodium phosphate was placed in the container, how many grams of sodium sulfate would be produce?
7. For the given reaction (unbalanced),

$$Ca(NO_3)_2 + Na_3PO_4 \rightarrow Ca_3(PO_4)_2 + NaNO_3$$

   a. how many grams of sodium nitrate are produced from 0.35 grams of sodium phosphate?
   b. how many grams of calcium nitrate are required to produce 5.5 grams of calcium phosphate?
8. For the given reaction (unbalanced),

$$Na_2S + Al(NO_3)_3 \rightarrow NaNO_3 + Al_2S_3$$

   a. how many grams of aluminum sulfide are produced from 3.25 grams of aluminum nitrate?
   b. how many grams of sodium sulfide are required to produce 18.25 grams of aluminum sulfide?

17.3 Limiting Reactant

Key Concepts

In this lesson, students learn to identify and calculated limiting and excess reactants in chemical reactions.

Lesson Objectives

- Identify the limiting reactant in a chemical reaction.
- Identify excess reactants in chemical reactions.
- Calculate the limiting reactant using the mole-mole ratios.
- Calculate the products using the limiting reactant and the mass-mass ratios.

Lesson Vocabulary

Limiting reactant The reactant that is completely consumed when a reaction is run to completion.

Excess reactant The reactant or reactants that are left over when all of the limiting reactant has been consumed.

Strategies to Engage

- Place five plates, four forks, four knives, and three spoons on a table in the classroom. Ask students how many place settings they can make from the given materials. Explain to students that because there are only three spoons, they can make only three place settings. Have students read the lesson introduction and compare the concept of limiting, and excess, reagents to the demonstration.
- Have students give examples of limiting reagents in everyday life such as making sandwiches.

Strategies to Explore

- Students often have trouble recognizing limiting reagent problems. Explain to students that any time they are given the amount of more than one reactant, they must first determine which reactant is the limiting reagent.
- Use the mathematical calculations in this lesson to reduce the reliance on language skills. As you go through each example problem, use them to explain the concepts explored in this lesson. DI **English Language Learners**

Strategies to Extend and Evaluate

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Limiting Reactant**. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Review Questions

Have students answer the lesson 17.3 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Consider the balanced reaction: $2\ \text{Al} + 6\ \text{HBr} \rightarrow 2\ AlBr_3 + 3\ H_2$.
   a. When 3.22 moles of Al reacts with 4.96 moles of HBr, how many moles of H_2 are formed?
   b. What is the limiting reactant?
   c. For the reactant in excess, how many moles are left over at the end of the reaction?

   1. Write the balanced equation for this reaction: copper (II) chloride reacts with sodium nitrate to form copper (II) nitrate and sodium chloride.
   2. If 15.0 grams of copper (II) chloride react with 20.0 grams of sodium nitrate, how much sodium chloride can be formed?
   3. What is the limiting reactant for this reaction?
   4. How much of the non-limiting reactant will be left over in this reaction?

17.4 Percent Yield

Key Concepts

In this lesson, students learn to calculate theoretical, actual, and percent yield of a chemical reaction.

Lesson Objectives

- Define theoretical and actual yield.
- Explain the difference between theoretical and actual yield.
- Calculate percent yield (reaction efficiency).

Lesson Vocabulary

theoretical yield The amount obtained when all of the limiting reactant has reacted in the balanced chemical equation.

actual yield The actual amount that is obtained from the experiment, and is always less than the theoretical yield.

percent yield $\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$

yield efficiency The percent yield of the reaction compared to the optimal yield.

Strategies to Engage

- Point out to students that if a chemical reaction occurs, in theory you can calculate how much of the product is created. However, in the real world, often not all the possible products are produced in a chemical reaction. Tell students that in this lesson they will explore reactions in which the product from a chemical reaction is less than was expected based on the balanced chemical equation.

Strategies to Explore

- Have students write a paragraph comparing and contrasting theoretical and actual yields.
- Use the mathematical calculations in this lesson to reduce the reliance on language skills. As you go through each example problem, use them to explain the concepts explored in this lesson. DI **English Language Learners**

Strategies to Extend and Evaluate

- On the board or chart paper, have students write a class summary of this lesson. Have one student come up with the first sentence and have students contribute sentences until the entire lesson has been summarized.

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Percent Yield**. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Review Questions

Have students answer the lesson 17.4 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Is it possible for the actual yield to be more than the theoretical yield?
2. What happens when competing reactions occur when performing an experiment in the lab?
3. If the actual yield is 4.5 g but the theoretical yield was 5.5 g, what is the percent yield for this data?
4. Solid aluminum and sulfur come together in a reaction to produce 7.5 g of aluminum sulfide. If 5.00 g of each solid react together, what is the percent yield? Remember to balance the reaction. $Al + S \rightarrow Al_2S_3$
   a. 32.0%
   b. 53.4%
   c. 96.2%
   d. 100.0%
5. In her experiment, Gerry finds she has obtained 3.65 g of lead(II) iodide. She knows her reaction was lead (II) nitrate that reacted completely with potassium iodide to produce lead (II) iodide and potassium nitrate. The potassium iodide produced is a brilliant yellow colored precipitate.

$$Pb(NO_3)_2(aq) + KI(aq) \rightarrow PbI_2(s) + KNO_3(aq)$$

   Gerry began with 5.00 g of potassium iodide. What was her percent yield? Remember to balance your equation first!
   a. 26.3%
   b. 36.0%
   c. 52.6%
   d. 72.0%
6. If a percentage yield in the reaction below was found to be 78.3%and the actual yield was 1.01 g, what was the original mass of the limiting reagent, oxygen.

$$2S + 3O_2 \rightarrow 2SO_3$$

   a. 0.515 g
   b. 0.773 g
   c. 1.01 g

d. 1.29 g

7. Bromine pentafluoride can be produced from a reaction between liquid bromine and fluorine gas. If 3.25 g of fluorine reacts with 2.74 g of bromine to produce 4.83 g of bromine pentafluoride, what is the percent yield of the product?
8. Ammonia can react with nitrogen in a reaction that is similar to a combustion reaction. The products, however, are nitrogen monoxide and water rather than carbon dioxide and water. In the reaction between ammonia and oxygen, 15 g of each reactant are placed in a container and 10.5 g of nitrogen monoxide was produced. What is the percent yield of the nitrogen monoxide?

17.5 Energy Calculations

Key Concepts

In this lesson, students explore energy changes in chemical processes.

Lesson Objectives

- Define endothermic and exothermic reactions in terms of energy and ΔH.
- Distinguish between endothermic and exothermic chemical changes.
- Write ΔH reactions for a given number of moles of reactants or products.

Lesson Vocabulary

law of conservation of energy The energy of the universe is constant and is therefore conserved.

potential energy Energy of position.

kinetic energy Energy of motion.

endothermic Energy is absorbed in the reaction, $\triangle H$ is positive or $\triangle H > 0$.

exothermic Energy is released in the reaction, $\triangle H$ is negative or $\triangle H < 0$.

heat of reaction, The change in energy from the products to the reactants $(\triangle H_{reaction} = H_{products} - H_{reactants})$.

enthalpy A measure of the energy content of a system.

Heat of Formation; the energy change when 1 *mole* of a substance is produced from its elements in their standard states.

Heat of combustion; the energy change that occurs when 1 *mole* of a fuel is reacted with oxygen.

Strategies to Engage

- Place 20 *ml* of water into each of two 50 *ml* beakers. Measure and record the temperature of the water in each beaker. Into one beaker, add 2g of baking soda. Into the other beaker, add 2g of sodium hydroxide. Measure

and record the temperature of each solution. The temperature of the sodium hydroxide solution should have increased, while the temperature of the baking soda solution should have decreased. Explain to students that when chemical processes take place, they are often accompanied by energy changes. Tell students that in this lesson they will explore these energy changes that occur in chemical processes.

Strategies to Explore

- It is important to define the system and surroundings. Point out to students that the reaction mixture constitutes the system. The surroundings are everything else.
- Perform the ***Exothermic Reaction*** *and* **Gummi Bear Oxidation (Exothermic Reaction)** demonstrations. These demonstrations are located in the Supplemental Lab Book.

Strategies to Extend and Evaluate

- Have students work in pairs to create a concept map relating the concepts explored in this lesson. Encourage students to include the drawings as well as the text in their concept map.
- Have students research examples of endothermic and exothermic processes that occur in everyday. Students should be prepared to discuss their findings with the rest of the class.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the lesson 17.5 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. How does the energy change of the bonds breaking compare to that of them forming? What happens with endothermic and exothermic reactions in terms of the energy?
2. The Bombardier beetle defends itself by squeezing fluid from the inner compartment in its body to the outer compartment. When the fluid reaches the outer compartment, an exothermic reaction takes place using the enzymes present in the beetle's body to produce $C_6H_4O_2$, the defense spray compound. The reaction that takes place in the outer body is shown below. Is this reaction endothermic or exothermic?

$$C_6H_4(OH)_2(aq) + H_2O_2(aq) \rightarrow C_6H_4O_2(aq) + 2H_2O(l) \qquad \triangle H = -203.6 \text{ kJ/mol}$$

3. Ozone, $O_3(g)$, is in the news quite frequently due to the environmental concerns about the hole in the stratosphere. The ozone present in the stratosphere protects us from the Sun's harmful ultraviolet rays. Which of the following choices best expresses the equation below?

$$3O_2(g) \rightarrow 2O_3(g) \qquad \triangle H = 286 \text{ kJ}$$

a. $\triangle H = 143$ kJ/mol and the reaction is endothermic
b. $\triangle H = 143$ kJ/mol and the reaction is exothermic
c. $\triangle H = 286$ kJ/mol and the reaction is endothermic
d. $\triangle H = 286$ kJ/mol and the reaction is exothermic

4. What can you say about the reaction below?

$$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g) \qquad \triangle H = -24.8 \text{ kJ}$$

a. it is endothermic and heat is absorbed
b. it is exothermic and heat is absorbed
c. it is endothermic and heat is released
d. it is exothermic and heat is released

5. For the reactions below, convert the thermochemical equations into the $\triangle H$ form and convert those in $\triangle H$ form into thermochemical equations.

a. $CH_4(g) + 2Cl_2(g) \rightarrow CH_2Cl_2(g) + 2HCl(g) + 230$ kJ
b. $Fe(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Fe^{2+}(aq) \qquad \triangle H = -152.3$ kJ
c. $N_2(g) + O_2(g) + 91.3 \text{ kJ} \rightarrow NO(g)$
d. $CH_3CH_2OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l) \qquad \triangle H = 1367$ kJ
e. $CH_4(g) + 2Cl_2(g) \rightarrow CH_2Cl_2(g) + 2HCl(g) \quad \triangle H = -230$ kJ (exothermic)
f. $Fe(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Fe^{2+}(aq) \qquad \triangle H = -152.3$ kJ (exothermic)
g. $N_2(g) + O_2(g) + 91.3 \text{ kJ} \rightarrow NO(g) \quad \triangle H = +91.3$ kJ (endothermic)
h. $CH_3CH_2OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l) \qquad \triangle H = 1367$ kJ (exothermic)

6. Dodecane is the main ingredient in diesel fuel. It has a formula $C_{12}H_{26}(L)$. When it undergoes combustion, 7561 kJ of energy is released per mole of dodecane. Using your knowledge of chemical reactions, write the combustion reaction for dodecane in both thermochemical and ΔH forms.

CHAPTER 18 The Kinetic Molecular Theory TE

Chapter Outline

Unit 6 Kinetic Molecular Explanation and the States of Matter

Outline

This unit, *Kinetic Molecular Explanation and the States of Matter*, includes the following chapters that explore the properties of the states of matter in terms of the Kinetic Molecular Theory.

- **Chapter 18 The Kinetic Molecular Theory**
- **Chapter 19 The Liquid State**
- **Chapter 20 The Solid State**

Overview

The Kinetic Molecular Theory

This chapter describes the molecular structure and properties of gases, and develops both the combined gas law and the universal gas law. The stoichiometry of reactions involving gases is also covered.

The Liquid State

This chapter covers the causes of the liquid condensed phase and the properties of liquids. It includes a section on the energy involved in liquid to gas phase changes and a section introducing colligative properties.

The Solid State

The various intermolecular forces of attraction are discussed in this chapter and the properties of solids produced by each type of intermolecular force of attraction are pointed out.

The Kinetic Molecular Theory

Key Concepts

The chapter *The Kinetic Molecular Theory* consists of seven lessons that describe the molecular structure and properties of gases, and develops both the combined gas law and the universal gas law. The stoichiometry of reactions involving gases is also covered.

- **Lesson 18.1 The Three States of Matter**
- **Lesson 18.2 Gases**
- **Lesson 18.3 Gases and Pressure**
- **Lesson 18.4 Gas Laws**
- **Lesson 18.5 Universal Gas Law**
- **Lesson 18.6 Molar Volume**
- **Lesson 18.7 Stoichiometry Involving Gases**

Overview

In these lessons, student will explore:

- The differences among the three states of matter.
- The behavior and properties of gases.
- The definition and measurement of gas pressure.
- Mathematical relationships among gas pressure, temperature, and volume.
- Calculations involving the universal gas law.
- The volume of a mole of gas at STP.
- Stoichiometric relationships involving reacting gas volumes.

Science Background Information

Elastic versus Inelastic Collisions

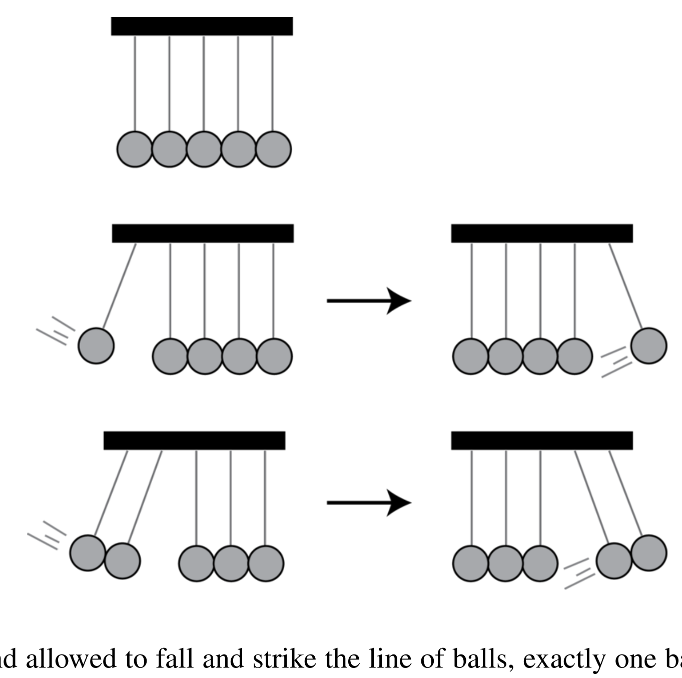
The momentum, (rho), of an object is defined as the mass of the object multiplied by its velocity, mv. The velocity of an object and the momentum of the object are vectors. That is a statement of either the velocity or the momentum of an object includes the direction that the object is traveling. The direction is an integral part of the measurement. If we assign the direction north to be positive direction, then a 5.0 kg object traveling north at 7.0 $meters/second$ will have a momentum of $+35\ kg \cdot m/s$. In this same system, a 5.0 kg object traveling south at 7.0 m/s will have a momentum of $-35\ kg \cdot m/s$. During collisions between objects, momentum is always conserved. In this system (consisting of these two objects), the total momentum of the system is $0\ kg \cdot m/s$ because $(+35\ kg \cdot m/s) + (-35\ kg \cdot m/s) = 0$. If these two objects collide and bounce directly backwards with velocity exactly opposite to their original velocities, the object that had a momentum of $+35\ kg \cdot m/s$ will now have a momentum of $-35\ kg \cdot m/s$ and the object whose original momentum was $-35\ kg \cdot m/s$ will now have a momentum of $+35\ kg - m/s$. The total momentum of the system is still $0\ kg \cdot m/s$ and momentum has been conserved (as it always is). If these two objects collide and stick

together (like two balls of Play Doh), both velocities become zero. In such a case, the momentum of each object is zero, the total momentum of the system is zero, and once again, momentum is conserved.

The kinetic energy, KE, of an object is defined as one-half the mass of an object multiplied by its velocity squared, $KE = \frac{1}{2}mv^2$. The kinetic energy of an object is NOT a vector. An 5.0 kg object traveling north at 7.0 m/s will have a $KE = \frac{1}{2}\ (5.0\ kg)(7.0\ m/s)^2 = 120\ kg \cdot\ m/s^2 = 120\ Joules$. In this same system, a 5.0 kg object traveling south at 7.0 m/s will also have a kinetic energy of 120 $Joules\ldots$ there is no direction associated with kinetic energy. Therefore, the total kinetic energy in this system is 240 $Joules\ldots$ the opposite directions of the ball's motions do not cause cancellation when dealing with kinetic energy. If these two objects collide and bounce directly backwards with velocities exactly opposite to their original velocities, each object will have the same kinetic energy it had before the collision and the total kinetic energy of the system will still be 240 $Joules\ldots$ kinetic energy has been conserved. If these two objects collide and stick together (like two balls of Play Doh), both velocities become zero. In such a case, the kinetic energy of each object is zero, the total kinetic energy of the system is zero, and kinetic energy is not conserved. Since energy, all forms considered, is conserved in all interactions except nuclear, the kinetic energy that was lost in the collision must be found in some other form, usually heat and sound. Make sure you understand that energy is conserved in non-nuclear interactions, but the form of the energy is not necessarily conserved. Specifically, KE is not conserved in the collision but energy in all forms is conserved. Mechanical energy may become electrical or electrical energy may become light, but when all forms of energy are added up, energy is conserved.

Considering collisions of all sorts, collisions between automobiles, collisions between tennis balls and walls, collisions between billiard balls, momentum is always conserved and kinetic energy is almost never conserved. When automobiles collide, metal parts are bent, causing parts to rub against each other, and friction turns kinetic energy into heat and sound. Even a tennis ball bouncing on the ground slowly loses energy of motion as it bounces. When the tennis ball strikes the ground, it is deformed and this deformation stores energy in the ball and as the ball regains its shape, the ball bounces back up in the air and the stored energy again becomes energy of motion. But, in the process, the deformation of the ball causes internal friction which is converted to heat and the ball will not bounce as high after each bounce. The tennis ball will bounce lower and lower until all the energy has been converted to heat. You probably cannot detect the temperature increase in a tennis ball but the same thing occurs when a hammer pounds on a nail and if you touch the nail after several strikes, you will feel the higher temperature.

Only a few collisions in nature come close to conserving kinetic energy. The collisions between billiard balls or between polished steel balls come quite close to conserving kinetic energy. A popular demonstration of conservation of momentum and conservation of kinetic energy features several polished steel balls hung in a straight line in contact with each other.



If one ball is pulled back and allowed to fall and strike the line of balls, exactly one ball will fly out the other side.

The other balls, including the one which was dropped will remain motionless.

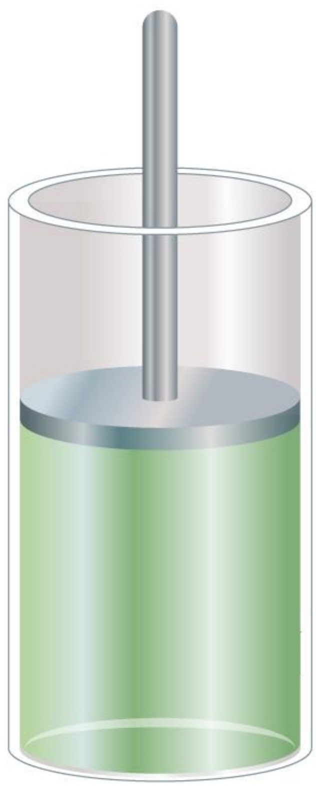
If two balls are pulled back and allowed to fall and strike the line of balls, exactly two balls will fly out the other side. The other balls, including the two that were dropped will remain motionless.

In the extreme case, if four balls are pulled back and allowed to fall, to strike the single motionless ball, four balls will fly out the other side, leaving one motionless ball.

The reason this strange phenomena occurs is that both momentum and kinetic energy are conserved in these collisions. Momentum would be conserved if one ball dropped at velocity X and two balls flew out the other side with velocity $\frac{1}{2}X$ but this would not conserve kinetic energy. In order for kinetic energy to be conserved, the same number of balls must fly out with the same velocity as the balls that were dropped.

When kinetic energy is conserved in a collision, physicists refer to the collision as a perfectly elastic collision. Why do we offer all this physics information to a chemistry teacher? The answer is that collisions between ideal gas particles are perfectly elastic collisions, that is, kinetic energy is conserved in collisions between gas particles. That's why when gas particles are bouncing around inside a container and exerting pressure, they do not gradually lose kinetic energy resulting in a lower pressure (as they would do if they were tennis balls).

There are cases, however, when gas particles do gain or lose kinetic energy without heat being added or removed from an external source. Consider a gas trapped in a closed cylinder fitted with a piston.



Consider the situation when the piston is held in position by hand. If the pressure inside the cylinder is greater than the external pressure, then releasing the piston will allow the gas insider the cylinder to push the piston higher. Moving the piston higher requires energy. Since the piston is being pushed higher by the gas inside the cylinder, the energy must come from the gas. The molecules of gas that strike the piston and push it upward are doing work (force $\times$ distance) and will lose some kinetic energy. Therefore, those molecules slow down. The average kinetic energy of the molecules in the cylinder becomes less and therefore, the temperature will be lower (temperature is proportional to the average kinetic energy of the molecules). Thus, the expansion of the gas against a force (outside pressure, gravity, etc.) causes the gas to cool slightly. Conversely, if you push the piston down, thus compressing the gas, your hand is doing work on the molecules the piston strikes. Those molecules will gain kinetic energy and so

the average *KE* of the molecules increases. The temperature of the gas will increase slightly. Thus, the compression of the gas raises its temperature slightly.

Suppose you have two boxes, one containing a gas and one containing a vacuum, and you open a valve between the boxes so the gas can expand into the vacuum. In this case, the gas is not pushing against anything, so it is not doing work and there will be no temperature change.

Pacing the Lessons

Use the table below as a guide for the time required to teach the lessons of *The Kinetic Molecular Theory*.

TABLE 18.1: 60 Minute Class Periods per Lesson

Lesson	Number of Class Periods
18.1 *The Three States of Matter*	0.5
18.2 *Gases*	0.5
18.3 *Gases and Pressure*	1.0
18.4 *Gas Laws*	2.0
18.5 *Universal Gas Law*	2.0
18.6 *Molar Volume*	0.5
18.7 *Stoichiometry Involving Gases*	1.5

Managing Materials

The following items are needed to teach the strategies and activities described in the Teachers Edition of the FlexBook for *The Kinetic Molecular Theory*.

TABLE 18.2: The Kinetic Molecular Theory Materials List

Lesson	Strategy or Activity	Materials Needed
18.1		
18.2	Engagement Activity	Bottle of perfume
18.3	Engagement Activity	Balloon
18.4	Engagement Activity	Aluminum soda can, hot plate, tongs, bucket.
18.5		
18.6	Engagement Activity	Poster board, tape.
18.7		

Multimedia Resources

You may find these additional web based resources helpful when teaching *The Kinetic Molecular Theory*:

- Lesson on temperature and absolute zero: http://www.colorado.edu/UCB/AcademicAffairs/ArtsSciences/physics/PhysicsInitiative/Physics2000/bec/temperature.html
- Particle motion computer simulation: http://intro.chem.okstate.edu/1314F00/Laboratory/GLP.htm

- Animated Boyle's Law: http://www.grc.nasa.gov/WWW/K-12/airplane/aboyle.html

Possible Misconceptions

Identify: Students may think that air does not have mass and does not take up space.

Clarify: Air is a mixture of gases. It contains 78% nitrogen, 21% oxygen, less than 1% argon, with trace amounts of other gases.

Promote Understanding: Tell students to wave their hands in front of their faces. Explain to student that although they cannot see the air, they could feel its effects. Measure and record the mass of a Ziploc bag. Fill the bag with air, and measure and record the mass again. Students will notice that air does have mass.

Making the FlexBook Flexible

An important advantage of the FlexBook is the ability it gives you, the teacher, to select the chapters and lessons that you think are most important for your own classes. You should also consult the standards correlation table that follows when selecting chapters and lessons to include in the FlexBook for your classes.

TABLE 18.3: **Standard Addressed by the Lessons in The Kinetic Molecular Theory**

Lesson	California Standards	NSES Standards	AAAS Benchmarks
18.1			
18.2	4a, 4f, 4g, 7a		
18.3	4a		
18.4	4c, 4d		
18.5	4c, 4h		
18.6	4b, 4i		
18.7			

18.1 The Three States of Matter

Key Concepts

In this lesson, students will explore the differences among the three states of matter in terms of both properties and structure.

Lesson Objectives

- The students will describe molecular arrangment differences among solids, liquids, and gases.
- The students will describe the basic characteristic differences among solids, liquids, and gases.

Lesson Vocabulary

phase Any of the forms or states, solid, liquid, gas, or plasma, in which matter can exist, depending on temperature and pressure.

kinetic The term "kinetic" refers to the motion of material bodies and the forces associated with them.

molecule In the kinetic theory of gases, any gaseous particle regardless of composition

Strategies to Engage

- Have each student draw a model of the particle arrangement in solids, liquids, and gases. Use this as an opportunity to clear up any misconceptions students may have about three states of matter.

Strategies to Explore

- Have student play a game of charades. Groups of students will act out one of the assumptions of the kinetic molecular theory, while the rest of the class tries to guess which assumption they are demonstrating.
- This lesson includes descriptions of the characteristics of solids, liquids, and gases. Before reading, prepare less proficient readers by having students write the following on the top of separate sheets of notebook paper:

Characteristics of Solids

Characteristics of Liquids

Characteristics of Gases

As they read each section have them write key points under each heading. This will give the students a quick reference and help them to organize the information. Instruct students to write a one-paragraph summary of the information they have read in each section. DI **Less Proficient Readers**

- Perform the ***Brownian Motion*** *and* **Molecular Motion/Kinetic Energy** demonstrations. These demonstrations are located in the Supplemental Lab Book.

Strategies to Extend and Evaluate

- Have students write a letter convincing the reader of the kinetic molecular theory. Instruct students to include real life examples in their letters.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 18.1 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Automobile brakes have a hose full of liquid connected to your brake pedal on one end and to the brake pads on the other end. When you press on the brake pedal, the force is transferred through the liquid and presses the brake pads against the wheels to slow or stop them. Brakes that use liquid in this fashion are called hydraulic brakes and the liquid is called hydraulic brake fluid. What don't they use air in the brake lines instead of liquid?
2. Why would it cause a problem if some air got into your liquid-filled brake lines?
3. If you had a 250 mL container full of helium gas and you transferred all of the gas into an empty 1.00 liter container, would the 1.00 liter container be full or only $1/4$ full? If something seems odd about this question, what do you think it is?

18.2 Gases

Key Concepts

In this lesson, students will explore the behavior and properties of gases.

Lesson Objectives

- The students will describe the relationship between molecular motion and Kelvin temperature.
- The students will describe random motion of gas molecules and explain how their collisions with surfaces cause pressure on the surface.
- The students will state that zero kinetic energy of molecules corresponds to 0 K and that there is no lower temperature.

Lesson Vocabulary

kinetic energy Kinetic energy is the energy a body possesses due to it motion, $KE = \frac{1}{2}mv^2$.

Kelvin temperature The absolute temperature scale where 0 K is the theoretical absence of all thermal energy (no molecular motion).

Strategies to Engage

- Open a bottle of perfume in the front of the room. Ask student to raise their hands when they are able to smell the scent. Ask a volunteer to explain, in terms of the kinetic molecular theory, why they are able to smell the scent.

Strategies to Explore

- Tell students to find the average low and high temperature of the previous day. Then have them convert the temperature in Fahrenheit into degrees Celsius and Kelvin. Students can draw three thermometers showing the equivalent temperatures on the three scales.
- Have less proficient readers make a main ideas/details chart as they read the lesson. Instruct them to divide a sheet of paper down the middle and record the main ideas on the left side and the details for each main idea on the right side. Have students save their chart for reviewing lesson content. DI **Less Proficient Readers**

Strategies to Extend and Evaluate

- Read each statement in the lesson summary. Have students indicate whether or not they understand each statement by using thumbs up or thumbs down to show "Yes" or "No". Whenever a student uses a thumbs down to show "No", use this as an opportunity to review this concept with the class. DI **English Language Learners**

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 18.2 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Ball A has a mass of 4 daltons and a speed of 16 meters per second. Ball B has a mass of 16 daltons. What velocity is necessary for ball B to have the same kinetic energy as ball A?
2. Suppose you blow up a balloon, tie off the opening, and place the balloon in a freezer for one hour. When you take the balloon out of the freezer, what will the most significant difference be in its appearance from when you put it in the freezer? What do you think will happen to the this difference as the balloon sits out in the room for a while?
3. Suppose you drive home from school on a hot day and when you get home, you check the pressure in your automobile tires. You find the tire pressure is over the manufacturer's recommended pressure and so you let some air out of the tires until the pressure is appropriate. What will the tire pressure be in the morning when you go out to go to school?
4. Weather balloons are large balloons that are used to carry meteorological instruments up through the atmosphere and radio back measurements on temperature, pressure, humidity, etc. as it passes through many different altitudes. When these balloons are filled with helium before they are released from earth, they are only filled a little more than 10% full. This provides enough lift to carry the instruments but the balloon would have more lift if it were filled completely. Why don't they fill the weather balloons to maximum capacity?

18.3 Gases and Pressure

Key Concepts

In this lesson, students will learn the definition of pressure and methods of measuring gas pressure.

Lesson Objectives

- The student will define pressure.
- The student will convert requested pressure units.
- The student will read barometers and both open-end and closed-end manometers.
- The student will apply the gas laws to relationships between the pressure, temperature, and volume of a gas.
- The student will state standard conditions for gases.

Lesson Vocabulary

barometer A barometer is an instrument used to measure atmospheric pressure.

manometer A manometer is a liquid column pressure measuring device.

Strategies to Engage

- Blow up a balloon. Ask students to list factors that influence the pressure of the air inside of the balloon. Students should respond that gas pressure is influenced by: the number of moles of gas in the container; its volume; and its temperature. Explain to students that in this chapter they will explore the relationships among these factors.

Strategies to Explore

- Ask students to look at Figure 18.5. Have them write a paragraph to describe what is happening in the illustration.
- Have students create a Venn diagram comparing and contrasting barometers and manometers.
- Explain to students that the force exerted on the floor when you stand on one foot is the same amount of force you exert on the floor when you stand on two feet. However, when you stand on one foot, the pressure or force per unit area is more.
- Perform the **'Magdeburn Hemispheres** demonstration. This demonstration is located in the Supplemental Lab Book.

Strategies to Extend and Evaluate

- Encourage interested students to research the science of scuba diving. Students should be prepared to share their findings with the rest of the class.

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Kinetic Molecular Theory"'. Ask students to complete the worksheet alone or in pairs as a review of lesson content.**

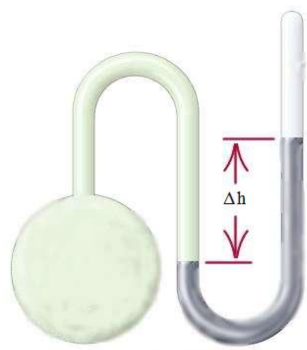
Review Questions

Have the students answers the Lesson 18.3 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

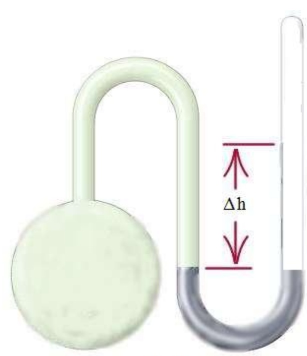
- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. The manometer shown is an closed-end manometer filled with mercury. If the atmospheric pressure in the room is 760. mm of Hg and $\triangle h$ is 65 mm of Hg, what is the pressure in the flask?



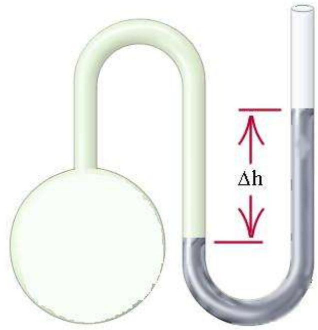
Closed-end Manometer

2. The manometer shown is an closed-end manometer filled with mercury. If the atmospheric pressure in the room is 750. mm of Hg and $\triangle h$ is 0 mm of Hg, what is the pressure in the flask?



Closed-end Manometer

3. The manometer shown is an open-end manometer filled with mercury. If the atmospheric pressure in the room is 750. mm of Hg and $\triangle h$ is 65 mm of Hg, what is the pressure in the flask?



4. The manometer shown is an open-end manometer filled with mercury. If the atmospheric pressure in the room is 760. mm of Hg and $\triangle h$ is 0 mm of Hg, what is the pressure in the flask?

Open-end Manometer

5. Explain why at constant volume, the pressure of a gas decreases by half when its Kelvin temperature is reduced by half.

18.4 Gas Laws

Key Concepts

In this lesson, the students will study mathematical relationships among gas pressure, temperature, and volume.

Lesson Objectives

- The students will state Boyle's Law, Charles' Law, and Gay-Lussac's Law.
- The students will solve problems using Boyle's Law, Charles' Law, and Gay-Lussac's Law.
- The students will state the combined gas law.
- Using the combined gas law, and given any five of the six variables, the students will solve for the sixth variable.

Lesson Vocabulary

barometer An instrument used to measure atmospheric pressure.

dalton The unified atomic mass unit, or Dalton, is a unit of mass used to express atomic and molecular masses. It is the approximate mass of a hydrogen atom, a proton, or a neutron. The precise definition is that it is one-twelfth of the mass of an unbound carbon-12 atom at rest.

manometer A liquid column pressure measuring device.

Strategies to Engage

- Place about 5 *ml* of water into an aluminum soda can and place it on a hot plate until you see steam rising from the can. Use a pair of tongs to grab the can and quickly invert the can into a bucket of cold water. Explain to students that the cold water quickly cooled the gas inside of the can. This pressure inside of the can decreased, and the atmospheric pressure was able to crush the can.

Strategies to Explore

- The internet is filled with simple demonstrations of the gas laws explored in this lesson. Ask groups of students to find, perform, and explain some of these demonstrations for their classmates.

- Use the mathematical calculations in this lesson to reduce the reliance on language skills. As you go through each example problem, use them to explain the concepts explored in this lesson. DI **English Language Learners**
- Divide students into groups of three or four to work on problems in this lesson. Assign one student in each group to serve as a reminder to the rest of the group members to use consistent units throughout the problems. Assign another student to serve as a reminder to convert temperatures from $^{\circ}C$ to K.
- Perform the ***Charles' Law with Balloons and Bunsen Burners*** *demonstration. This demonstration is located in the Supplemental Lab Book.*

Strategies to Extend and Evaluate

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Combined Gas Law**. Ask students to complete the worksheet alone or in pairs as a review of lesson content.

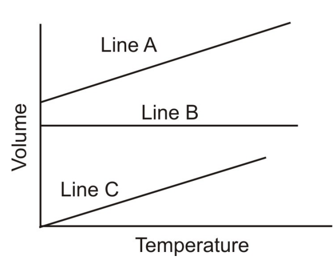
Review Questions

Have students answer the Lesson 18.4 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. When a sample of gas is placed in a larger container at the same temperature, what happens to the total FORCE of the molecules hitting the walls?
2. When a sample of gas is placed in a larger container at the same temperature, what happens to the pressure exerted by the gas?
3. If molecules of H_2 (mol.mass $= 2$), O_2(mol.mass $= 32$), and N_2(mol.mass $= 28$) are all placed in the same container at the same temperature, which molecules will have the greatest average kinetic energy?
4. If molecules of H_2(mol.mass $= 2$), O_2(mol.mass $= 32$), and N_2(mol.mass $= 28$) are all placed in the same container at the same temperature, which molecules will have the greatest velocity?
5. If X and Y are quantities that are related to each other by inverse proportion, what will the value of Y become when the value of X is increased by a factor of five?
6. Under what conditions will the value for the constant, K, change in the equation for Boyle's Law, $PV =$ K.
7. A sample of gas has a volume of 500. mL under a pressure of 500. mm of Hg. What will be the new volume of the gas if the pressure is reduced to 300. mm of Hg at constant temperature?



8. A graph is made illustrating Charles's Law. Which line in the picture at right would be appropriate assuming temperature is measured in Kelvin?

9. At constant pressure, the temperature of a sample of gas is decreased. Will the volume of the sample
    a. increase
    b. decrease
    c. remain the same?
10. A sample of gas has its temperature increased from $-43°C$ to $47°C$ at constant pressure. If its volume at $-43°C$ was 500. mL, what is its volume at $47°C$?
11. A gas is confined in a rigid container and exerts a pressure of 250. mm of Hg at a temperature of $17°C$. To what temperature must this gas be cooled in order for its pressure to become 216 mm of Hg? Express this temperature in $°C$.
12. What is the abbreviation used to indicate standard conditions of temperature and pressure?
13. A sample of gas has a volume of 500. mL at standard conditions. Find its volume at $47°C$ and 800. torr.
14. A sample of gas has a volume of 100. L at $17°C$ and 800. torr. To what temperature must the gas be cooled in order for its volume to become 50.0 L at a pressure of 600. torr?

18.5 Universal Gas Law

Key Concepts

In this lesson, students learn and practice calculations involving the universal gas law.

Lesson Objectives

- The students will solve problems using the Universal Gas Law, $PV = nRT$.
- The students will state Avogadro's Law of equal molecules in equal volumes under the same conditions of temperature and pressure.
- The students will calculate molar mass from $mm = \frac{gRT}{PV}$, given mass, temperature, pressure, and volume.

Lesson Vocabulary

universal gas law constant, R is a constant equal $\frac{PV}{nT}$ where the pressure, volume, moles, and temperature of the gas are P, V, n, and T, respectively. The value and units of R depend on the units of P and V. Commonly used values of R include;

$$82.055\ mL^3\ atm\ K^{-1}\ mol^{-1},\ 0.082055\ L\ atm\ K^{-1}\ mol^{-1},\ 8.314\ JK^{-1}\ mol^{-1},\ 8.314\ Pa\ m^3\ K^{-1}\ mol^{-1}.$$

Strategies to Engage

- Facilitate a discussion with students about the relationships among pressure, temperature, and volume of a gas explored so far in this chapter.

Strategies to Explore

- Use the mathematical calculations in this lesson to reduce the reliance on language skills. As you go through each example problem, use them to explain the concepts explored in this chapter. DI **English Language Learners**
- Divide students into groups of three or four to work on problems in this lesson. Assign one student in each group to serve as a reminder to the rest of the group members to use consistent units throughout the problems. Assign another student to serve as a reminder to convert temperatures from $^\circ C$ to K.

Strategies to Extend and Evaluate

- Encourage interested students to research the science of car engines and how they relate to the gas laws. Students should be prepared to share their findings with the rest of the class.
- Have students create study cards of the equations explored in this chapter.

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Universal Gas Law**. Ask students to complete the worksheet alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 18.5 Review Questions that are listed at the end of the lesson in the FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. What volume will 2.00 moles of hydrogen gas occupy at 2.62 atm of pressure and $300.^\circ C$?
2. How many moles of gas are required to fill a volume of 8.00 liters at 2.00 atm and 273 K?
3. What is the molar mass of a gas if its density is 1.30 g/L at STP?

18.6 Molar Volume

Key Concepts

In this lesson, students learn the volume of one mole of any gas at Standard Temperature and Pressure (STP) and applications of that information.

Lesson Objectives

- The students will apply the relationship 1.00 *mole* of any gas at standard conditions will occupy 22.4 *L*.
- The students will convert gas volume at STP to moles and to molecules and vice versa.
- The students will apply Dalton's Law of Partial Pressures to describe the composition of a mixture of gases.

Lesson Vocabulary

diffusion The movement of particles from areas of higher concentration to areas of lower concentration of that particle.

partial pressure The pressure that one component of a mixture of gases would exert if it were alone in a container.

molar volume The volume occupied by one mole of a substance in the form of a solid, liquid, or gas.

Strategies to Engage

- Perform the ***Rate of Diffusion at Different Tempertures*** *demonstration. This demonstration is located in the Supplemental Lab Book.*

Strategies to Explore

- Use the mathematical calculations in this lesson to reduce the reliance on language skills. As you go through each example problem, use them to explain the concepts explored in this chapter. DI **English Language Learners**

Strategies to Extend and Evaluate

- Ask groups of students to use poster board and tape to build a cube that will hold exactly 1.00 *mol* of a gas at STP. Award a prize to the first group who is able to build the cube with the correct dimensions and explain the calculations involved.

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Molar Volume and Partial Pressure**. Ask students to complete the worksheet alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 18.6 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. A 1.00 L container of helium gas at 1.00 atm pressure and a 1.00 L container of hydrogen gas at 2.00 atm are both transferred into a 1.00 L container containing nitrogen gas at 3.00 atm. What is the final pressure in the final container holding all three gases (assuming no temperature change)?
2. For the situation described in problem #1, what will be the partial pressure of the helium in the final container?
3. What conditions of temperature and pressure cause gases to deviate from ideal gas behavior?
4. At STP, how many molecules are in 89.6 liters of gas?
5. If 1.00 liter of gas A at STP and 1.00 liter of gas B at STP are both placed into a 2.00 liter evacuated container at STP, what will the pressure be in the 2.00 liter container?
6. Consider the gases CO and N_2. Which of the following will be nearly identical for the two gases at $25°C$ and 1.0 atm?
   a. I only
   b. III only
   c. I and II only
   d. II and III only
   e. I, II, and III
      a. average molecular speed
      b. rate of effusion through a pinhole
      c. density
7. The density of an unknown gas at 2.0 atm and $25°C$ is determined to be 3.11 g/L. Which of the following gases is the unknown most likely to be?
   a. CH_4
   b. F_2
   c. N_2O_4
   d. O_2
   e. CF_2Cl_2

18.7 Stoichiometry Involving Gases

Key Concepts

In this lesson, students explore stoichiometric relationships involving gas volume.

Lesson Objectives

- The students will solve stoichiometry problems involving converting gas volume at STP to moles and vice versa.
- The students will solve stoichiometry problems involving gas volume to gas volume under any conditions of temperature and pressure.

Strategies to Engage

- Explain to students that in the last lesson they explored molar volume of gases at STP. In this lesson they will solve stoichiometry problems involving volume relationships in balanced equations. Have students read the introduction, lesson vocabulary and lesson objectives, then facilitate a discussion with students about how they think they will perform these calculations.

Strategies to Explore

- Point out to students the importance of writing the correct units throughout the problem, and that in the end, all of the units must cancel except for the desired unit. This will prevent students from bypassing the mole ratio and attempting to convert the volume of the given gas directly to the volume of the desired gas.

Strategies to Extend and Evaluate

- Have students complete the lab ***Finding the Molar Mass of a Gas Experimentally***. This lab is located in the Supplemental Lab Book.

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Stoichiometry Involving Gases**. Ask students to complete the worksheet alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 18.7 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. How many liters of hydrogen gas are required to react with 25.0 L of nitrogen gas according to the following equation? $N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$
2. How many grams of ammonia will be formed from 25.0 L of nitrogen gas measured at STP according to the equation in question 29?

Chapter 19 The Liquid State TE

Chapter Outline

The Liquid State

Outline

The chapter *The Liquid State* consists of five lessons that cover the causes of the liquid condensed phase and the properties of liquids. It includes a section on the energy involved in liquid to gas phase changes and a section introducing colligative properties.

- **Lesson 19.1 The Properties of Liquids**
- **Lesson 19.2 Forces of Attraction**
- **Lesson 19.3 Vapor Pressure**
- **Lesson 19.4 Boiling Point**
- **Lesson 19.5 Heat of Vaporization**

Overview

In these lessons, students will explore:

- The behavior of liquids.
- Intermolecular forces of attraction.
- Vaporization, condensation, and vapor pressure.
- The relationship between vapor pressure and boiling point.
- The energy changes involved in cooling and heating.

Science Background Information

This information is provided for teachers who are just beginning to instruct in this subject area.

Refrigeration

Vaporization is the phase change from liquid to gas at the boiling point of the liquid. When this phase change occurs below the boiling point of the liquid, it is called evaporation. Liquids undergo evaporation because while the average temperature of its molecules is less than the boiling point, some of the molecules have temperatures above the boiling point. These hot molecules are the ones that leave the liquid phase and enter the gaseous phase. During both vaporization and evaporation, the amount of liquid that leaves the liquid phase and enters the gaseous phase absorb the heat of vaporization. When the ambient temperature of a gaseous substance is above the boiling point of the liquid of that substance, scientists called the substance a gas. But when the ambient temperature of a gaseous substance is below the boiling point of the liquid of that substance, they call it a vapor. Hence, gaseous water at an ambient temperature of $120^{\circ}C$ is water gas and gaseous water at an ambient temperature of $70^{\circ}C$ is water vapor.

The process of evaporation has long been used to cool food and drink.

Canteens are frequently covered in fabric or carried in a fabric holder. The canteen user wets the fabric when filling the canteen so that as the water evaporates from the fabric, it absorbs the heat of vaporization from the canteen and cools the canteen, making the water more pleasant to drink.

Some people put butter on the dinner table with the dish holding the butter sitting inside another dish half-filled with water. As the water in the outside dish evaporates, it absorbs the heat of vaporization, cools the butter dish and keeps the butter from melting.

Before the days of the portable ice chest, people who took bottled or canned drinks on a picnic would often keep the drinks in a fabric bag that they would soak with water on arrival at the picnic spot. The evaporation of the water would keep the drinks much cooler than if they were sitting out on a table.

The function of refrigerators and air conditioners also involve the heat of vaporization of liquids.

Many gaseous substances can be compressed until they become liquids. That is, the molecules are pushed together forcefully until they touch and the gas becomes a liquid. In this process, the gas also gives up the heat of vaporization as it becomes a liquid. By compressing the coolant to a liquid outside the refrigerator, the phase change gives up the heat of vaporization outside the refrigerator. The liquid coolant is then pumped through a tube inside the refrigerator where it is allowed to vaporize back to gas, thus absorbing the heat of vaporization inside the refrigerator. Then the gas is pumped outside the refrigerator and again compressed to liquid, giving up the heat of vaporization. In this manner, heat is absorbed from inside the refrigerator and given off outside the refrigerator. The inside gets colder and the outside gets warmer. This is why you can feel heat coming from the back or from underneath a refrigerator. This is also why the compressor for an air conditioner must be outside the house. It wouldn't do much good to absorb the heat and release the heat both inside the house. So much for the idea of cooling the house by leaving the refrigerator door open.

Pacing the Lessons

Use the table below as a guide for the time required to teach the lessons of *The Liquid State*.

TABLE 19.1: 60 Minute Class Periods per Lesson

Lesson	Number of Class Periods
19.1 *The Properties of Liquids*	0.5
19.2 *Forces of Attraction*	1.0
19.3 *Vapor Pressure*	0.5
19.4 *Boiling Point*	0.5
19.5 *Heat of Vaporization*	1.0

Managing Materials

The following items are needed to teach the strategies and activities described in the Teachers Edition of the FlexBook for *The Liquid State*.

TABLE 19.2: The Liquid State Materials List

Lesson	Strategy or Activity	Materials Needed
19.1		
19.2	Exploration Activity	Pennies, dropper pipets, alcohol, distilled water
19.3		
19.4		
19.5	Exploration Activity	Ice, beakers, thermometer, hot plate, ring stand assembly

Multimedia Resources

You may find these additional web based resources helpful when teaching *The Liquid State*:

- Fill-in-the-blank worksheet generator: http://www.theteacherscorner.net/printable-worksheets/make-your-own/fill-in-the-blank/

Possible Misconceptions

Identify: Students may think that boiling and vaporization have the same meaning. Also, students may not understand the difference between evaporation and boiling.

Clarify: Vaporization is the transition of a liquid to a gas. Vaporization can take place in two ways: evaporation and boiling. Evaporation occurs when some particles within a liquid have more energy than others, and are able to escape from the surface of the liquid as gas or vapor. Evaporation takes place below the boiling point of the liquid. Boiling happens when the vapor pressure of the liquid is equal to atmospheric pressure.

Promote Understanding: Have students construct a simple concept map illustrating the relationships among boiling, vaporization, and evaporation. The concept map should show that evaporation and boiling are types of vaporization.

Identify: Students may think that the temperature of a liquid increases as it boils.

Clarify: The temperature of a boiling liquid never goes above its boiling point no matter how much heat is applied to it.

Promote Understanding: Have students measure and record the temperature of a sample of water every at 30 *second* interval as it boils. Ask students to construct a graph of time vs. temperature. Students should see that the temperature does not change.

Making the FlexBook Flexible

An important advantage of the FlexBook is the ability it gives you, the teacher, to select the chapters and lessons that you think are most important for your own classes. You should also consult the standards correlation table that follows when selecting chapters and lessons to include in the FlexBook for your classes.

TABLE 19.3: Standard Addressed by the Lessons in The Liquid State

Lesson	California Standards	NSES Standards	AAAS Benchmarks
Lesson 19.1			
Lesson 19.2	2d, 2h		
Lesson 19.3	7c		
Lesson 19.4	7c		
Lesson 19.5	7c, 7d		

19.1 The Properties of Liquids

Key Concepts

In this lesson, students explore the behavior and properties of liquids.

Lesson Objectives

- The student will explain the basic behavior and characteristics of liquids using the molecule arrangement present in liquids.

Lesson Vocabulary

incompressible The terms compressibility and incompressibility describe the ability of molecules in a fluid to be compacted (made more dense).

Strategies to Engage

- Ask students what they already know about liquids. Use this opportunity to gauge student understanding of the properties of liquids and to clear up any misconceptions.

Strategies to Explore

This lesson includes a description of the basic behavior and properties of liquids. Before reading, prepare less proficient readers by having students write the following on the top of separate sheets of notebook paper:

- Liquids Maintain Their Volume But Take the Shape of Their Container
- Liquids Have Greater Densities Than Gases
- Liquids are Almost Incompressible
- Liquids Diffuse More Slowly Than Gases

As students read each section, have them write key points under each heading. This will give the students a quick reference and help them to organize the information. Instruct students to write a one-paragraph summary of the information they have read in each section. DI **Less Proficient Readers**

Strategies to Extend and Evaluate

- Encourage interested students to research fluids and write a paragraph explaining why liquids and gases are classified as fluids.
- Have students work in pairs or teams to write a poem about liquids. Their poems should explain what liquids are, some of their properties, and how they differ from gases.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 19.1 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. The molar volumes of solid silicon and solid bromine under the same conditions are different and the molar volumes of liquid silicon and liquid bromine under the same conditions are different, but the molar volumes of gaseous silicon and gaseous bromine under the same conditions are exactly the same. Explain.

19.2 Forces of Attraction

Key Concepts

In this lesson, students explore intermolecular forces of attraction.

Lesson Objectives

- The student will identify liquids whose intermolecular forces of attraction are due to London dispersion forces, polar attractions, and hydrogen bonding.
- The student will describe some of the unique properties of water that are due to hydrogen bonding.
- The student will select from comparative compounds, the ones most likely to form hydrogen bonding.
- The student will select from comparative compounds whose intermolecular forces are London dispersion forces, the one most likely to have the strongest intermolecular forces.

Lesson Vocabulary

hydrogen bond The exceptionally strong polar attraction between a hydrogen atom in one molecule and a highly electronegative atom *(N, O, F)* in another molecule.

London dispersion forces Electrostatic attractions of molecule or atoms for nearby atoms or molecules caused by the temporary unsymmetrical distribution of electrons in electron clouds.

Strategies to Engage

- Prior to beginning this lesson, have students look up examples of terms that begin with the prefixes "intra" and "inter". Ask them to write down the meanings of the words. Facilitate a discussion with students about how these prefixes relate to molecules.

Strategies to Explore

- Have students write a paragraph comparing and contrasting intermolecular forces and chemical bonds.
- Have students place three drops of distilled water and three drops of alcohol on two separate pennies. Ask students write a paragraph explaining their observations in terms of intermolecular forces of attraction in each liquid.
- Ask students to look at Figure 19.3 and write a paragraph describing what is happening in the illustration.

Strategies to Extend and Evaluate

- Encourage interested groups of students to create Keynote or PowerPoint slideshow presentations explaining the intermolecular forces of attraction explored in this lesson to share with the rest of the class. Students should include illustrations and examples of each type of intermolecular force.

Lesson Worksheets

There are no worksheets for this lesson.

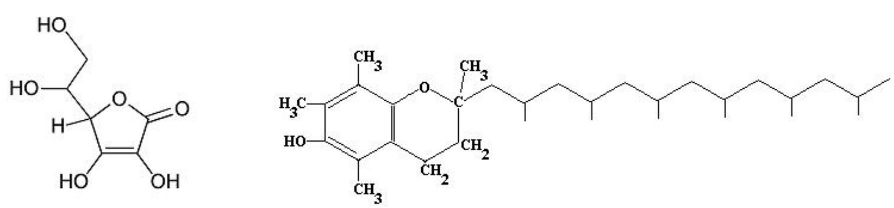
Review Questions

Have students answer the Lesson 19.2 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Which of the following molecules would you expect to have the higher melting point? Why?
   a. CH_4 or H_2S
   b. H_2O or H_2S
   c. HCl or Cl_2
   d. NaI or NH_3
   e. SF_4 or CH_4
2. The structures of vitamins C and E are shown above. Which of the following statements is correct?



Vitamin C Vitamin E

   a. Vitamin E has more opportunities for hydrogen bonds than vitamin C.
   b. The melting point of vitamin E is likely to be higher than that of vitamin C.
   c. Vitamin C is likely to be very soluble in a non-polar solvent.
   d. Vitamin C should have a higher solubility in water than vitamin E.
   e. Vitamin C would be described as a "fat-soluble" vitamin.

19.3 Vapor Pressure

Key Concepts

In this lesson, students will learn about vaporization, condensation, and vapor pressure.

Lesson Objectives

- The students will describe the processes of evaporation and condensation.
- The students state the factors that control the rates of evaporation and condensation.
- The students will describe vapor pressure equilibrium.

Lesson Vocabulary

condensation The process whereby a gas or vapor is changed to a liquid.

equilibrium vapor pressure The pressure that is exerted, at a given temperature, by the vapor of a solid or liquid in equilibrium with the vapor.

evaporation The escape of molecules from a liquid into the gaseous state at a temperature below the boiling point.

heat of condensation The quantity of heat released when a unit mass of a vapor, condenses to liquid at constant temperature.

heat of vaporization The quantity of heat required to vaporize a unit mass of liquid at constant temperature.

vapor The gaseous phase of a substance that exists even though the temperature is below the boiling point of the substance.

Strategies to Engage

- Have students read the review questions at the end of this section. This way, students will be familiar with the types of information that they will explore in this section.

Strategies to Explore

- Have students create a graph of the vapor pressure of water at various temperatures from the table shown in this lesson. Ask them to correctly identify the boiling point of water on the graph, and then write a paragraph explaining the graph in their own words.
- Ask students to look at Figure 19.9 and write a paragraph to describe what is happening in the illustration.

Strategies to Extend and Evaluate

- Have students create a concept map of vocabulary terms in this lesson. Tell students to relate vocabulary terms to the concepts explored in this lesson, and to correctly illustrate the relationships between the terms and the concepts.
- Have students use grid paper to make a crossword puzzle using the vocabulary terms. Ask students to exchange papers with a classmate and solve each other's puzzles.
- Challenge students to write an illustrated children's story that includes examples of condensation and vaporization they encounter in an average day (water puddles evaporating, fog forming on mirrors).
- Have each student record the four sentences in this section that most clearly represent the main ideas. Read key sentences in the text and have students raise their hands if they have recorded that sentence. Facilitate a discussion in which students defend their selections.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 19.3 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. A flask half-filled with water is sealed with a stopper. The space above the water contains hydrogen gas and water vapor in vapor pressure equilibrium with the liquid water. The total pressure of the two gases is 780. mm of Hg at $20.^{\circ}C$. The vapor pressure of water at $20.^{\circ}C$ is 19 mm of Hg. What is the partial pressure of the hydrogen gas in the flask?
2. Describe all the reasons that the remaining liquid cools as evaporation occurs.
3. Describe all the reasons that the remaining gas gets hotter as condensation occurs.
4. The apparatus above can be used to determine the vapor pressure of benzene. With a vacuum in the top of the tube, the mercury rises to the height shown. When a small amount of liquid benzene is injected into the space at the top of the tube, it floats on the mercury. The benzene will evaporate and eventually reach vapor pressure equilibrium. The mercury in the tube will pushed down further by the pressure of the benzene vapor in the tube. Neglecting the effect of the liquid benzene, what would be the calculated vapor pressure of benzene?
5. Water vapor and hydrogen gas are sealed in a cylinder fitted with a piston at $60^{\circ}C$. The partial pressure of the hydrogen gas is 0.35 atm and the vapor pressure of the water is 0.20 atm at this temperature. The total

pressure in the cylinder is 0.55 atm. If the piston is pushed down until the volume is half the original volume, what will be the pressure in the cylinder?

19.4 Boiling Point

Key Concepts

In this lesson, students will learn the relationship between vapor pressure and boiling point.

Lesson Objectives

- The students will state the relationship between boiling point, vapor pressure, and ambient pressure.
- Given a vapor pressure table for water, and the ambient pressure, the students will determine the boiling point of water for specified conditions.

Lesson Vocabulary

boiling point The temperature at which the vapor pressure of a liquid equals the surrounding pressure.

normal boiling point The temperature at which the vapor pressure of a liquid equals 1.00 atmosphere.

Strategies to Engage

- Ask students to look at Figure 19.11. Facilitate a discussion with students about why they think the water is able to boil at $20^\circ C$. Explain to students that in this lesson they will explore how and why boiling point changes with changes in pressure.

Strategies to Explore

- Have students create a Venn diagram comparing and contrasting boiling point with normal boiling point.
- Perform the **'Boiling Water in a Paper Cup** demonstration. This demonstration is located in the Supplemental Lab Book.
- Have students complete the lab ***The Race to*** 110 *Degrees**C***. This lab is located in the Supplemental Lab Book.

Strategies to Extend and Evaluate

- Have each student write five fill-in-the-blank statements with the blank at the end of the sentence about key concepts explored in this lesson. Have students exchange papers with another student who will try to complete

the sentence by filling in the blank. Have them hand the papers back to the original student who will assign a grade. Encourage students to discuss any incorrect answers. Students can also generate fill-in-the-blank worksheets at: http://www.theteacherscorner.net/printable-worksheets/make-your-own/fill-in-the-blank/

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 19.4 Review Questions that are listed at the end of the lesson in their FlexBook.

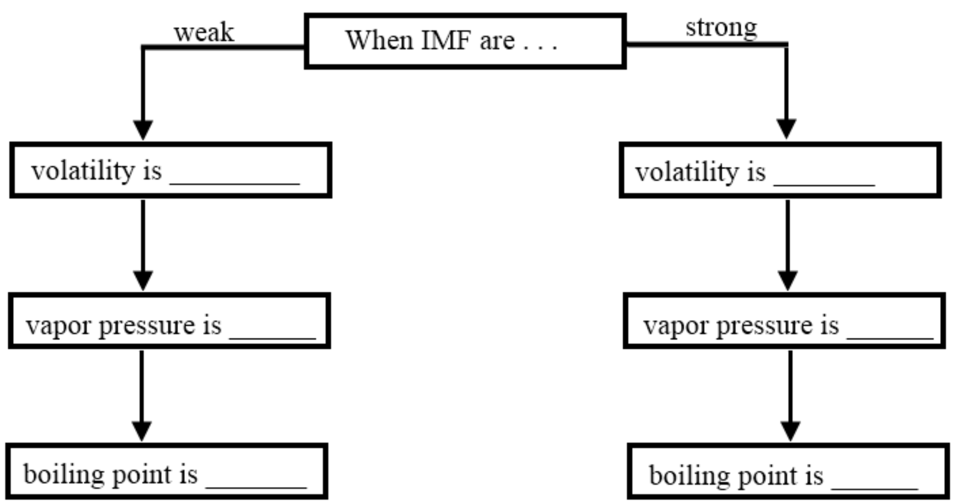
Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. What happens to the boiling point of a liquid if the pressure exerted on the surface of the liquid is increased?

2. How can you make water boil without heating it?

3. Fill in the diagram with either "high" or "low" to show how intermolecular forces of attraction influence the **volatility**, **vapor pressure**, and **boiling point** of a substance.



19.5 Heat of Vaporization

Key Concepts

In this lesson, students will learn to calculate the energy changes during phase changes.

Lesson Objectives

- The student will calculate energy changes during phase changes.
- The student will explain the slopes of various parts of heating and cooling curves.

Lesson Vocabulary

heat of condensation The quantity of heat released when a unit mass of a vapor condenses to liquid at constant temperature.

heat of vaporization The quantity of heat required to vaporize a unit mass of liquid at constant temperature.

Strategies to Engage

Strategies to Explore

- Have groups of students create a heating curve for water. Ask students to write a materials list, procedure, and data table and approve them before proceeding. Students should then write a paragraph that includes the vocabulary terms explaining the heating curve.
- Use the mathematical calculations in this lesson to reduce the reliance on language skills. As you go through each example problem, use them to explain the concepts explored in this lesson. DI **English Language Learners**
- Use Figure 19.12 to explain as many concepts as possible. Relate concepts such as heat of vaporization, heat of condensation, and specific heat to Figure 18. DI **English Language Learners**

Strategies to Extend and Evaluate

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 19.5 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. How much heat is required to vaporize 200. grams of water at $100.^\circ C$ and 1.00 atm pressure? $\triangle H_{\text{VAP}}$ for water is 2.25 kJ/g.
2. How much heat is required to raise 80.0 grams of water from $0^\circ C$ to $100.^\circ C$ with no phase change occurring? The specific heat of water is $4.18 \text{ J/g.}^\circ C$.

CHAPTER 20 The Solid State-HSC TE

Chapter Outline

The Solid State

Outline

The chapter *The Solid State* consists of four lessons that discuss the various intermolecular forces of attraction. Also discussed are the properties of solids produced by each type of intermolecular force of attraction.

- **Lesson 20.1 The Molecular Arrangement in Solids Controls Solid Characteristics**
- **Lesson 20.2 Melting**
- **Lesson 20.3 Types of Forces of Attraction for Solids**
- **Lesson 20.4 Phase Diagrams**

Overview

In these lessons, students will explore:

- The characteristics of solids.
- Energy changes that occur when a substance melts.
- Forces of attraction within solids.
- The reading and interpretation of phase diagrams.

Science Background Information

This information is provided for teachers who are just beginning to instruct in this subject area.

Liquid Crystals

LCD or liquid crystal displays have become a ubiquitous part of our technology landscape. Now appearing as computer and television screens and other electronic displays, even in new automobile dashboard devices. Yet for

chemistry students with an alert ear, the seemingly contradictory term Liquid Crystal, should at very least, merit additional questions. We know that crystalline lattices are structures characteristic of the solid phase, with atoms or ions limited in their positions to vibrational motion, in place of the translational capabilities due to the decrease in intermolecular attractive forces in the liquid state.

There are currently thousands of different compounds, however that display behaviors intermediate between that of the liquid and solid state. Liquid crystals are arranged in a regular, orderly pattern yet their individual molecules can flow like liquids. The types of molecules that tend to form liquid crystals are usually cylindrically – shaped, with a polar group at one end of the molecule. This shape allows different opportunities for arrangement, such as orientation in the same directions (nematic), or alignment in layers (smectic). These arrangements are due to the presence of dipole – dipole or hydrogen bonding interactions or, at times, a combination of both forces.

FIGURE 20.1

Due to their unusual structural arrangements, liquid crystals exhibit interesting thermal, optical, and electronic properties. Some liquid crystal samples will react to changes in temperature. You may have used a body thermometer that display the temperature with a liquid crystal. Pressure–sensitive liquid crystals have been implemented in the design of fingerprint detection devices. More commonly, the application of an electric or magnetic field can result in the realignment of a liquid crystal sample which in turn causes a change in the visual display. Most nematic liquid crystals are transparent or translucent but with the application of an electrical field, the molecular orientation alters and the display becomes opaque.

The current popularity of devices containing liquid crystals continues to grow as the demand for light–weight, flexible display technology increases. Some future uses for this technology include the incorporation of liquid crystals into carbon nanotubes to create three–dimensional arrays. Another interesting potential application of liquid crystals includes their use in an anti-cancer drug, as well as the use of liquid crystals in cosmetics and personal care products.

Pacing the Lessons

Use the table below as a guide for the time required to teach the lessons of *The Solid State*.

FIGURE 20.2

TABLE 20.1: 60 Minute Class Periods per Lesson

Lesson	Number of Class Periods
20.1 *Molecular Arrangement in Solids Controls Solid Characteristics*	0.5
20.2 *Melting*	1.0
20.3 *Types of Forces of Attraction for Solids*	0.5
20.4 *Phase Diagrams*	0.5

Managing Materials

The following items are needed to teach the strategies and activities described in the Teachers Edition of the FlexBook for *The Solid State*.

TABLE 20.2: The Solid State Materials List

Lesson	Strategy or Activity	Materials Needed
20.1	Engagement Activity	Plastic sandwich bags and cornstarch
20.2		
20.3		
20.4		

Multimedia Resources

You may find these additional web based resources helpful when teaching *The Solid State*:

- Lesson on Characteristics of Solids, Liquids, and Gases: http://www.chem.purdue.edu/gchelp/liquids/character.html
- Phase Diagram Learning Activity: http://www.wisc-online.com/objects/index_tj.asp?objID=GCH6304

TABLE 20.3: (continued)

Lesson 20.4	California Standards	NSES Standards	AAAS Benchmarks

20.1 The Molecular Arrangement in Solids Controls Solid Characteristics

Key Concepts

In this lesson, students explore the characteristics of solids.

Lesson Objectives

- The students will describe the molecular arrangement in solids.
- The students will use the molecular arrangement in solids to explain the incompressibility of solids.
- The students will use the molecular arrangement in solids to explain the low rate of diffusion in solids.
- The students will use the molecular arrangement in solids to explain the ability of solids to maintain their shape and volume.

Lesson Vocabulary

Strategies to Engage

- Give each student a plastic sandwich bag that contains one cup of cornstarch. Ask students to add one cup of water to the solid and knead the material in the bag for three minutes. Draw four columns on the board. Have each student tell whether they think the material is a solid, liquid, either, or neither, and place their reasoning in the appropriate column.

Strategies to Explore

- Have students write down the lesson objectives, leaving about 5 or 6 lines of space in between. As you explore the lesson, have students write the "answer" to each objective.

Strategies to Extend and Evaluate

- Encourage interested students to research liquid crystal technology. Students should be prepared to share their findings with their classmates.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 20.1 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Fill in the types of phase changes left blank in the chart below.

Solid $\rightarrow$ Liquid_____________
Liquid $\rightarrow$ Gas_____________
Solid $\rightarrow$ Gas___ Sublimation ___
Liquid $\rightarrow$ Solid _____________
Gas $\rightarrow$ Liquid_____________
Gas $\rightarrow$ Solid___Deposition___

20.2 Melting

Key Concepts

In this lesson, students explore energy changes that occur when a substance melts.

Lesson Objectives

- The students will explain why it is necessary for a solid to absorb heat during melting even though no temperature change is occurring.
- Given appropriate thermodynamic data, the students will calculate the heat required to raise temperatures of a given substance with no phase change.
- Given appropriate thermodynamic data, the students will calculate the heat required to melt specific samples of solids with no temperature change.
- Given appropriate thermodynamic data, the students will calculate the heat required to produce both a phase change and a temperature change, for a given sample of solid.

Lesson Vocabulary

crystal A solid consisting of plane faces and having definite shape with the atoms arranged in a repeating pattern.

freezing The phase change from liquid to solid.

freezing point The temperature at which a liquid changes to a solid.

fusion

1. The change of a liquid to a solid.
2. A nuclear reaction in which two or more smaller nuclei combine to form a single nucleus.

heat of condensation The quantity of heat released when a unit mass of vapor condenses to a liquid at constant temperature.

heat of fusion The quantity of heat released when a unit mass of liquid freezes to a solid at a constant temperature.

heat of vaporization The quantity of heat absorbed when a unit mass of liquid vaporizes to a gas at constant temperature.

joule A basic unit of energy in the SI system, equal to one Newton-meter.

melting The phase change from solid to liquid.

melting point The temperature at which a substance changes from the solid phase to the liquid phase.

Strategies to Engage

- Ask students what they already know about melting. Use this opportunity to gauge student understanding of the melting process and to clear up any misconceptions.
- Have students read the review questions at the end of this section. This way, students will be familiar with the types of information that they will explore in this section.

Strategies to Explore

- Use the mathematical calculations in this lesson to reduce the reliance on language skills. As you go through each example problem, use them to explain the concepts explored in this chapter. DI **English Language Learners**

Strategies to Extend and Evaluate

- Have students create a concept map of the lesson vocabulary terms. Tell students to relate vocabulary terms to the concepts explored in this lesson, and to correctly illustrate the relationships between the terms and the concepts.

Lesson Worksheets

Copy and distribute the lesson worksheets in the *CK-12 Chemistry Workbook* titled **Heat Transfer** and **Calorimetry**. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 20.2 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

Use the thermodynamic data given in the **Table** 20.4 to answer problems 1 - 5.

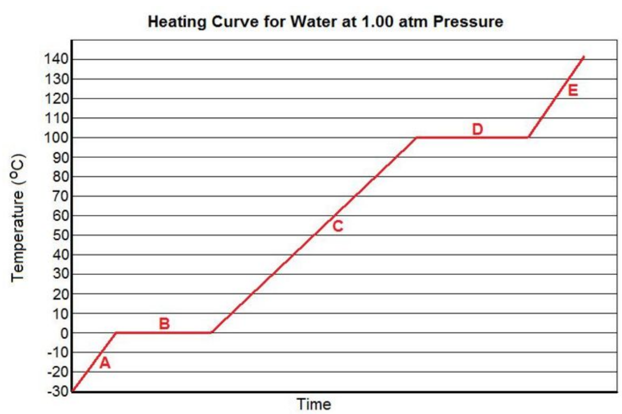
TABLE 20.4: Thermodynamic Data of Various Substances

	Water	Cesium, Cs	Silver, Ag
Melting Point	$0°C$	$29°C$	$962°C$
Boiling Point	$100.°C$	$690.°C$	$2162°C$

TABLE 20.4: (continued)

	Water	Cesium, Cs	Silver, Ag
$\triangle H_{\text{fusion}}$	334 J/g	16.3 J/g	105 J/g
$\triangle H_{\text{vaporization}}$	2260 J/g	669 J/g	2362 J/g
Specific Heat, C, for Gas	$2.01\ J/g \cdot {}^\circ C$	$0.167\ J/g \cdot {}^\circ C$	$0.159\ J/g \cdot {}^\circ C$
Specific Heat, C, for Liquid	$4.18\ J/g \cdot {}^\circ C$	$0.209\ J/g \cdot {}^\circ C$	$0.294\ J/g \cdot {}^\circ C$
Specific Heat, C, for Solid	$2.09\ J/g \cdot {}^\circ C$	$0.251\ J/g \cdot {}^\circ C$	$0.235\ J/g \cdot {}^\circ C$

1. How many Joules are required to melt 100. grams of silver at its normal melting point with no temperature change?
2. How many Joules are required to boil 150. grams of cesium at its normal boiling point with no temperature change?
3. How many Joules are required to heat 200. g of liquid water from $25^\circ C$ to steam at $125^\circ C$ under normal pressure?
4. How many Joules are required raise the temperature of 1.00 gram of water from $-269^\circ C$ (the current temperature of space) to $1.60 \times 10^{15}\ {}^\circ C$ (the estimated temperature of space immediately after the big bang)?
5. How many Joules are required to raise the temperature of 1000. g of cesium from $-200.^\circ C$ to $+200.^\circ C$?
6. Why does the boiling point of water increase with increasing surrounding pressure?
7. Why must heat be absorbed to melt a solid even though both the solid and the liquid are at the same temperature?



8. What is happening to the water in section B?
9. What is happening to the water in section A?
10. Why are the slopes of the lines in sections A, C, and E different?

20.3 Types of Forces of Attraction for Solids

Key Concepts

In this lesson, students will learn the nature of the forces of attraction within solids.

Lesson Objectives

- The students will describe the metallic bond, and explain some of the solid characteristics that are due to metallic bonding.
- Given characteristics of a solid such as conductivity of solid and liquid phase, solubility in water, malleability, and so on, the students will identify the type of solid, i.e. the attractive forces holding the solid in solid form.

Lesson Vocabulary

alloy A substance composed of a mixture of two or more elements and having metallic properties.

conductivity The property of being able to transmit heat and/or electricity.

conductor A substance that can transmit heat and/or electricity.

ductility The property of a substance that allows it to be drawn into a wire.

electrical conductivity The ability of a substance to transmit an electric current.

malleable The property of being able to be hammered or rolled into sheets.

metallic bond The attractive force that binds metal atoms together. It is due to the attractive force that the mobile electrons exert on the positive ions.

specific heat The amount of energy necessary to raise 1.00 *gram* of a substance by $1.00°C$.

Strategies to Engage

Strategies to Explore

Ask students to look at Figure 20.3 and write a paragraph describing what is happening in the illustration.

This lesson includes a description of different types of solids. Before reading, prepare less proficient readers by having students write the following on the top of separate sheets of notebook paper:

- Ionic Solids
- Metallic Solids
- Network Solids
- Amorphous Solids

As they read each section have them write key points under each heading. This will give the students a quick reference and help them to organize the information. Instruct students to write a one-paragraph summary of the information they have read in each section. DI **Less Proficient Readers**

Strategies to Extend and Evaluate

- Have students work in pairs or teams to write a poem about solids. Their poems should explain what solids are and the types of forces of attraction for solids.
- Have students use grid paper to make a crossword puzzle using the vocabulary terms. Ask students to exchange papers with a classmate and solve each other's puzzles.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 20.3 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Identify the most important type of inter-particle force present in the following solids that is responsible for binding the particles into a solid.
   a. He
   b. NO
   c. HF
   d. $BaCl_2$
   e. CH_4
   f. $NaNO_3$
   g. CO_2
   h. $CHCl_3$
   i. pure Mg
   j. diamond
2. Predict which substance in the following pairs would have the stronger force of attraction between molecules and justify your answer.
   a. CO_2 or OCS

b. PF_3 or PF_5
c. NaI or I_2
d. H_2O or H_2S
e. solid argon or solid sodium
f. HF or HBr

3. In the following groups of substances, pick the one that has the requested property and justify your answer.
   a. highest boiling point: HCl, Ar, F_2
   b. highest melting point: $H_2O, NaCl, HF$
   c. lowest vapor pressure at $20°C$: Cl_2 , Br_2, I_2
4. An unknown solid is not soluble in water or CCl_4. The solid conducts electricity and has a melting point of $800°C$. Identify the most likely attractive forces holding the particles in the solid state.
5. An unknown solid is soluble in water but not in CCl_4. The solid does not conduct electricity but its liquid does. The solid shatters when hammered and has a melting point of $1430°C$. Identify the most likely attractive forces holding the particles in the solid state.
6. Why would you expect ionic solids to have higher melting points that polar solids?
7. Why does the melting point of water decrease with increasing surrounding pressure?
8. List the following substances in order of increasing boiling points: $BaCl_2$, H_2, CO, HF, Ne, CO_2.
   a. H_2
   b. CH_3OH
   c. CH_2Cl_2
   d. KCl
   e. CO

Select your answers for questions 9, 10, and 11 from these choices.

9. Which of these substance is most likely to be a solid at $25°C$ and 1.0 atm?
10. Which of these substances is capable of hydrogen bonding?
11. Which of the substances has its solid properties governed by London dispersion forces?
12. Place these molecules, CF_4, $CaCl_2$, and ICl, in order of decreasing melting points (highest first).
    a. $CF_4 > CaCl_2 > ICl$
    b. $CaCl_2 > ICl > CF_4$
    c. $CaCl_2 > CF_4 > ICl$
    d. $ICl > CF_4 > CaCl_2$
    e. $CF_4 > ICl > CaCl_2$

20.4 Phase Diagrams

Key Concepts

In this lesson, students will explore the reading and interpretation of phase diagrams.

Lesson Objectives

- The students will read specific requested information from a phase diagram.
- The students will state the primary difference between a generic phase diagram, and a phase diagram for water.

Lesson Vocabulary

critical pressure The pressure required to liquefy a gas at its critical temperature.

critical temperature The highest temperature at which it is possible to liquefy the substance by increasing pressure.

Strategies to Engage

- Explain to students that in this lesson, they will learn how to show the relationships among the solid, liquid, and vapor states of a substance in one simple diagram.

Strategies to Explore

- Use Figure 20.6 to explain as many concepts as possible. Relate concepts such as critical temperature and critical pressure to Figure 11. DI **English Language Learners**

Strategies to Extend and Evaluate

- On the board or chart paper, have students write a class summary of this lesson. Have one student come up with the first sentence and have students contribute sentences until the entire lesson has been summarized.

Lesson Worksheets

There are no worksheets for this lesson.

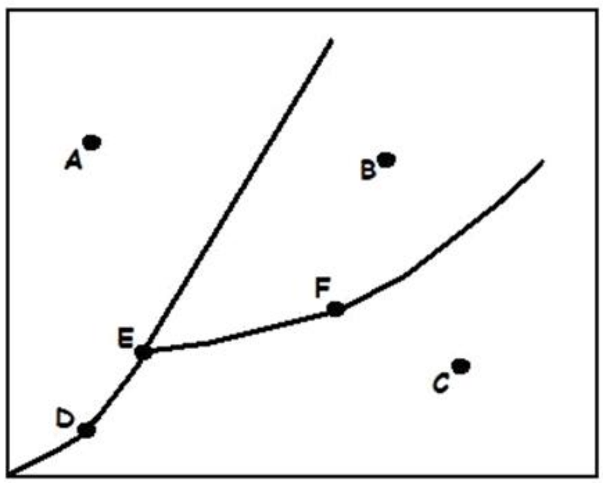
Review Questions

Have students answer the Lesson 20.4 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Consider the phase diagram below.



Name the phases that may be present at each lettered point in the diagram.

20.5 References

1. . http://commons.wikimedia.org/wiki/File:EOS_rear.jpg. CC-BY-SA
2. . http://commons.wikimedia.org/wiki/File:479563754_8ef9e978a7.jpg. Public Domain

CHAPTER 21 The Solution Process TE

Chapter Outline

Unit 7 Solutions and Their Behavior

Outline

This unit, *Solutions and Their Behavior*, includes two chapters that cover the solution process and the behavior of ions in solution.

- **Chapter 21 The Solution Process**
- **Chapter 22 Ions in Solution**

Overview

The Solution Process

This chapter describes solvation, concentration calculations, solubility, and colligative properties of solutions.

Ions in Solution

This chapter covers dissociation, electrolytes and non-electrolytes, reactions between ions in solution, and ionic and net-ionic equations.

The Solution Process

Outline

This chapter, *The Solution Process*, consists of nine lessons that cover solvation, concentration calculations, solubility, and colligative properties of solutions.

Overview

In these lessons, students will explore:

- The composition of solutions.
- The relationship between molecular structure and solution formation.
- Vocabulary associated with solutions.
- Methods of expressing solution concentration.
- The information provided by a solubility graph.
- The factors that affect the solubility of solids and gases.
- The colligative properties of solutions.
- The similarities and differences among solutions, colloids, and suspensions.
- Methods used to separate mixtures.

Science Background Information

This information is provided for teachers who are just beginning to instruct in this subject area.

Properties of Solutions

Definitions

solution: homogeneous mixture of a solute dissolved in a solvent

solute: component present in smaller amount

solvent: component present in greatest amount

The Solution Process

As a solute crystal is dropped into a solvent, the solvent molecules begin to attack and pull apart the solute ions or

molecules. Solvent molecules surround the solute molecules in a process called hydration, forming a solvent cage around the solute particles as the solute dissolves in the solvent.

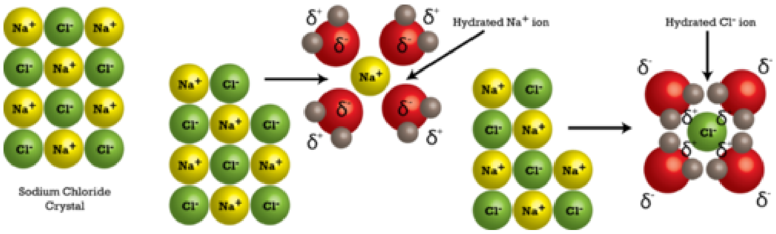


FIGURE 21.1

The illustration above shows the hydration process of ions. A similar process occurs when dissolving polar molecules that do not form ions. Each polar solute molecule is attached to one or more polar solvent molecules.

Energy Changes During Solution Formation

Three types of interactions must be considered for solution formation.

- solvent-solvent attraction
- solute-solute attraction
- solvent-solute attraction

In order for a solute to dissolve, the solvent-solute attractions must be at least equal to the solvent-solvent and solute-solute attractions.

So why don't all solids dissolve in liquids?

If solvent-solute interaction can't compete with solute-solute and solvent-solvent interactions, they remain separated.When solute-solute or solvent-solvent interactions are stronger than solute-solvent interactions, solute and solvent stay separated. When solute-solvent interactions are as strong as solute-solute and solvent-solvent interactions, solute and solvent mix.

Saturated Solutions and Solubility

Definitions

solubility: Maximum amount of solute dissolved in solvent at specific temp.

unsaturated: contains less than the maximum amount of solute that a solvent can hold at specific temperature.

saturated: contains the maximum amount of solute that a solvent can hold at specific temperature.

supersaturated: contains more than the maximum amount of solute that a solvent can hold at specific temperature.

How do supersaturated solutions form?

At higher temperatures, solvents can hold more solute than at lower temperatures. If a given amount of solute is dissolved in a solvent at a higher temperature, then allowed to cool without being disturbed, the solute will remain in solution. The solution is unstable, though, and the solute will crystallize rapidly if disturbed.

Factors Affecting Solubility

A general rule referred to as "Like Dissolves Like" applies to solution formation. Ionic compounds and polar molecules dissolve in polar solvents but do not dissolve in non-polar solvents. Non-polar solutes dissolve in non-polar solutes but do not dissolve in polar solvents. The reason was referred to earlier in the section on Energy Changes During Solution Formation. When polar molecules are introduced to non-polar solvents, the solute-solute attractions are greater than the solute-solvent attractions and the solute does not dissolve. When non-polar solutes

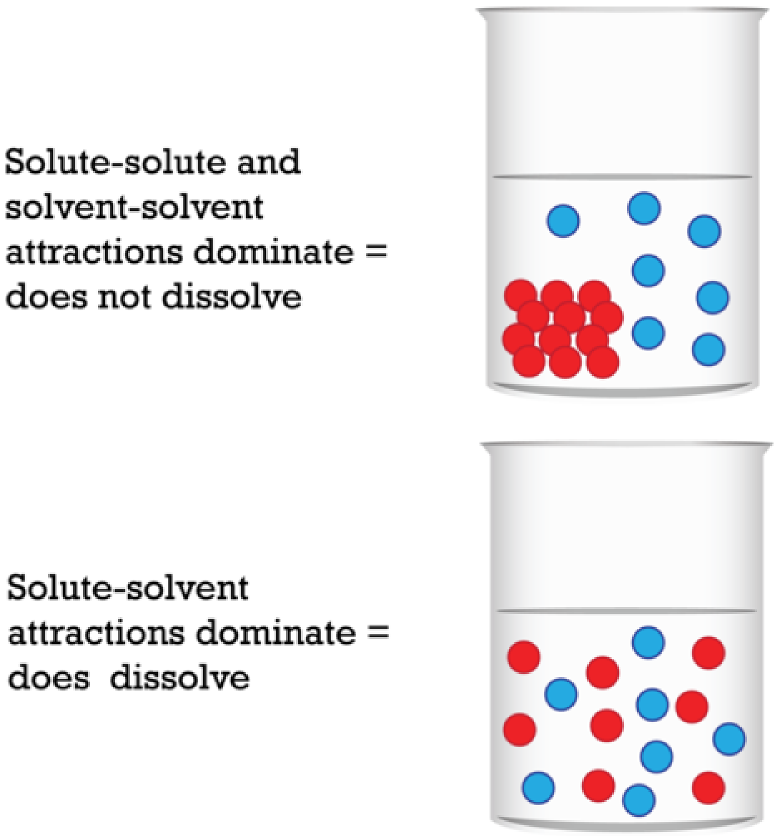


FIGURE 21.2

are introduced to polar solvents, the solvent-solvent attractions are greater than the solute-solvent attractions and the solute does not dissolve.

Liquid-Liquid Solutions

Polar molecules will mix (be miscible with) other polar molecules.

Non-polar molecules will mix (be miscible with) other non-polar molecules.

Polar molecules will not mix (be immiscible with) non-polar molecules.

Solid-Liquid Solutions

Ionic and polar molecular compounds dissolve in polar solvents. (This is a general rule but there are some ionic compounds that do not dissolve in polar solvents . . . for the final answer, you must check the solubility rules.)

Non-polar molecular compounds dissolve in non-polar solvent.

Molecules containing polar bonds and non-symmetrical shapes will be polar. Molecules without polar bonds or with polar bonds but in symmetrical shape will be non-polar.

Some solids do not dissolve in any solvent

- Network covalent solids (eg. graphite, quartz) do not dissolve in any solvent.
- Metals do not dissolve in any solvent. (They may react but they don't dissolve.)

For example, of the following substances, $NaCl$, CCl_4, NH_3, and $C_{diamond}$, will be soluble in water and which will be soluble in hexane (a non-polar liquid)?

$NaCl$ is ionic and will therefore, dissolve in water but not in hexane. CCl_4 is non-polar and will therefore, dissolve

in hexane but not in water. NH_3 is polar and will therefore, dissolve in water but not in hexane. $C_{diamond}$ is a network covalent solid and will not dissolve in water or hexane.

Gas-Liquid Solutions

Gas solubility and Pressure Effects:

Henry's Law: The solubility of gas is proportional to partial pressure of gas above liquid.

The solubility of a gas increases when the partial pressure of the gas above the liquid increases and solubility decreases when the partial pressure of the gas above the liquid decreases. This is because a higher partial pressure of gas above the liquid means that a larger number of gas molecules are in contact with the liquid surface and therefore, a larger number of gas molecules will enter the liquid phase (dissolve).

Example

In a container of carbonated beverage that is sealed, the carbonation will remain indefinitely. When the container is opened, the gas above the liquid escapes, lowering the partial pressure of gas above the liquid, and the carbonation (carbon dioxide gas) will gradually decrease as the carbon dioxide gas comes out of solution.

Gas solubility and Temperature:

As the temperature increases, the solubility of a gas in a liquid decreases (in most cases). Gases are more soluble in liquids at lower temperatures.

Example

When a pan of water is heated on the stove, gas bubbles form on the bottom and sides of the pan long before the water is nearing the boiling point. The bubbles are not bubbles of water vapor but rather are bubbles of air that have come out of solution as the water warmed.

Solid solubility and Temperature:

As the temperature increases, the solubility of a solid in a liquid increases (in most cases).

Example

More sugar will dissolve in hot tea than in iced tea.

Methods of Expressing Concentration

Mass percent of solute = $\frac{mass\ of\ solute}{mass\ of\ solution} \times 100$

5.0 $grams$ of salt dissolved in 495 $grams$ of water is a 1% solution of salt.

(ppm), parts per million = $\frac{milligrams\ of\ solute}{kilograms\ of\ solvent}$

5.0 $milligrams$ of salt dissolved in 5.0 kg of water is 1 ppm.

(ppb), parts per billion = $\frac{micrograms\ of\ solute(\mu g)}{kilograms\ of\ solvent}$

5.0 $milligrams$ (5000 $micrograms$) dissolved in 5.0 kg of water is 1000 ppb.

Mole Fraction (X) (has no units since ratio of two similar quantities) = $\frac{moles\ of\ solute}{total\ moles\ insolution}$

200. $grams$ of $CaBr_2$ (molar mass = 200. g/mol) dissolved in 500. $grams$ of water (molar mass = 18.0 g/mol) = 1.00 mol of $CaBr_2$ dissolved in 27.8 $mols$ of water so the mole fraction, $X_{CaBr_2} = \frac{1.00\ mol}{28.8\ mols} = 0.0347$ and $X_{H_2O} = \frac{27.8\ mols}{28.8\ mols} = 0.965$.

$$\text{Molarity (M)} = \frac{mols\ of\ solute}{liters\ of\ solution}$$

200. $grams$ of $CaBr_2$ dissolved in 500. mL of solution = $\frac{1.00\ mol}{0.500\ L} = 2.00\ M$

$$\text{Molality (m)} = \frac{mols\ of\ solute}{kg\ of\ solvent}$$

200. *grams* of $CaBr_2$ dissolved in 500. *grams* of water = $\frac{1.00\ mol}{0.500\ kg} = 2.00\ m$

For molarity dilution problems, use $M_1V_1 = M_2V_2$ where M = molarity and V = volume.

All concentration expressions are independent of temperature except molarity. Since the volume of solution changes with temperature, so does molarity.

Examples of converting from one concentration unit to another.

1. Calculate the molarity of a sodium chloride solution, which is 5.0% $NaCl$(molar mass $= 58.5\ g/mol$). The density of the solution is 1.03 g/mL.

Imagine a 100. mL sample of solution is taken. The mass of the total solution is 103 *grams*. 5.15 *grams* of the solution is $NaCl$ and 97.85 *grams* of the solution is water. The 5.15 g of $NaCl$ is 0.0880 mol. Therefore, the molarity will be $\frac{0.0880\ mol}{0.100\ L} = 0.880\ M$.

2. Calculate the molarity of a 0.500 m glucose (molar mass $= 180.\ g/mol$) solution if the density of the solution is 1.16 g/mL.

Imagine a sample that contains 1000. g of water. This sample will also contain 0.500 mol of glucose which has a mass of 90.0 *grams*. Therefore, the total mass of the sample is 1090. *grams*. Dividing this total mass by the density yields the volume of the sample, 940. mL. Therefore, the molarity will be $\frac{0.500\ mol}{0.940\ L} = 0.532\ M$.

Colligative Properties

colligative properties: properties that depend on the number of solute particles in solution and not on the nature of the solute particles

non-electrolytes: exist as molecules in solution (do not dissociate into ions)

electrolytes: exist as ions in solution

Lowering the Vapor Pressure (Non-electrolytes)

vapor pressure: pressure exerted by vapor in equilibrium with its liquid or solid

A substance that has very low vapor pressure is nonvolatile, whereas one that exhibits a vapor pressure is volatile.

Oil is considered nonvolatile while gasoline is volatile.

Adding a solute lowers the concentration of solvent molecules in liquid phase since solute particles on the surface of the solution block solvent molecules from evaporating.

FIGURE 21.3

Adding a solute to a solvent lowers the vapor pressure of the solvent. There are two suggested explanations for why the addition of a solute lowers the vapor pressure of a solution. Since they seem equally valid, both will be presented here. Remember that only the molecules on the surface of a liquid are able to evaporate.

In a pure solvent, all the molecules at the surface are solvent molecules. Therefore, the entire surface area is available for evaporation and the forces to be overcome are the attractive forces between the solvent molecules. One of the explanations says that in a solution, some of the surface molecules are solute molecules and since these solute molecules take up some of the surface area, less surface area is available for evaporation. Therefore, the rate of evaporation of the solvent will be lower and so the vapor pressure will be lower at the same temperature. The other explanation says that the attractive forces between the solvent molecules and the solute molecules are greater than the attractive forces between solvent molecules and therefore, the solvent molecules will not evaporate at as high a rate. Once again vapor pressure will be lowered. Both explanations start with the same premises and end with the same result so there doesn't seem to be a reason to choose between them.

Boiling-Point Elevation and Freezing Point Depression (Non-electrolytes)

Boiling-Point Elevation:

A liquid boils when its vapor pressure equals the surrounding (ambient) pressure. For example, pure water has a vapor pressure of 760 mm of Hg at $100^\circ C$. Therefore, when liquid water is raised to $100^\circ C$, it boils. If a non-volatile solute is added to water, the vapor pressure of the solution is lower than the vapor pressure of the pure solvent. Such a solution will have a vapor pressure less than 760 mm of Hg at $100^\circ C$ and therefore, will not boil at this temperature. In order for the vapor pressure of the solution to exhibit a vapor pressure of 760 mm of Hg, the temperature must be raised higher than $100^\circ C$. Therefore, the boiling point of the solution is greater than the boiling point of the pure solvent.

The increased boiling point is determined as follows: $T_b = T_{bp}water\ +\ \Delta T_b$, where $T_b\ =$ boiling point of solution, $T_{bp}water\ =$ b. p. of pure solvent, ΔT_b = change in b.p.

ΔT_b is calculated using $\Delta T_b = K_b m$, where $m =$ molal concentration of solute and $K_b\ =$ molal boiling point constant.

Example: Calculate the boiling point of a solution containing 1.25 mol of glucose in 0.250 kg of water using $K_b = 0.52^\circ C/m$.

$$\text{molality} = \frac{mols\ solute}{kg\ of\ solvent} = \frac{1.25\ mols}{0.250\ kg} = 5.0\ m$$
$$\Delta T_b = (5.0\ m)(0.52^\circ C/m) = 2.6^\circ C$$
$$\text{B.P. of solution} = 100.^\circ C + 2.6^\circ C = 102.6^\circ C$$

Freezing-Point Depression:

Two things happen when ice and water are placed in contact: molecules on the surface of the ice escape into the water (melting), and molecules of water are captured on the surface of the ice (freezing). When the rate of freezing is the same as the rate of melting, the amount of ice and the amount of water won't change on average. The ice and water are said to be in dynamic equilibrium with each other. The balance between freezing and melting can be maintained at $0^\circ C$, the melting point of water, unless conditions change in a way that favors one of the processes over the other.

The balance between freezing and melting processes can easily be upset. If the ice/water mixture is cooled, the molecules move slower. The slower-moving molecules are more easily captured by the ice, and freezing occurs at a greater rate than melting. Conversely, heating the mixture makes the molecules move faster on average, and melting is favored. Adding salt (or other non-volatile solute) to the system will also disrupt the equilibrium. Consider replacing some of the water molecules with molecules of some other substance. The foreign molecules dissolve in the water, but do not pack easily into the array of molecules in the solid. The total number of water molecules captured by the ice per second goes down, so the rate of freezing goes down. The rate of melting is unchanged by the presence of the foreign material, so melting occurs faster than freezing.

That's why salt melts ice. To re-establish equilibrium, you must cool the ice-saltwater mixture to below the usual melting point of water. For example, the freezing point of a 1 $MNaCl$ solution is roughly $-3.4^\circ C$. Solutions will

always have such a freezing point depression. The higher the concentration of salt, the greater the freezing point depression.

The new freezing point is determined as follows: $T_{fp} solution = T_f solvent - \Delta T_f$ where T_{fp} solution = freezing point of solution, T_f solvent = freezing point of pure solvent, and ΔT_f = freezing point depression.

ΔT_f is calculated using $\Delta T_f = K_f$ m, where $m =$ molal concentration of solute and $K_f =$ molal freezing point constant.

Example: Calculate the freezing point of a solution containing 1.25 mol of glucose in 0.250 kg of water using $K_f = 1.86°C/m$.

$$Molality = \frac{1.25\,mols}{0.250\,kg} = 5.00\,m$$
$$\Delta T_f = K_f m = (1.86°C/m)(5.00\,m) = 9.3°C$$
$$T_{fp}\ solution = 0.0°C - 9.3°C = -9.3°C$$

Osmotic Pressure (Non-electrolytes)

semipermeable membrane: allows solvent molecules to pass through but blocks the passage of solute molecules

osmosis: net movement of solvent molecules through semipermeable membrane from pure solvent or more dilute solution to more concentrated solution

osmotic pressure (π)*:* pressure required to stop osmosis

isotonic: when two solutions have equal osmotic pressure

hypertonic: the more concentrated of two solutions that are not isotonic

hypotonic: the less concentrated solution of two solutions that are not isotonic

We can calculate osmotic pressure (π) at a given temperature: $\pi = MRT$, where M $=$ molarity of solute, $R = 0.0821 L \cdot atm/mol \cdot K$, and $T =$ Kelvin temperature.

Example: A 0.125 M sample of seawater was taken at $25°C$. Calculate the osmotic pressure of the seawater sample.

$$\pi = MRT = (0.125\,mol/L)(0.0821\,L \cdot\ atm/mol \cdot\ K)(298\,K) = 3.06\,atm$$

Colligative Properties of Electrolyte Solutions

For colligative properties, electrolyte solutions have another factor to consider beyond those of non-electrolyte solutions. Colligative properties are controlled by the molality of the number of particles in solution. A 1.0 m solution of non-electrolyte will contain 1.0 mol of particles because the molecules do not dissociate in water solution. Electrolytes, on the other hand, do dissociate in water solution. In the case of an electrolyte like $NaCl$, a 1.0 m solution will contain 2.0 m solution of ions. If the electrolyte were $CaCl_2$, a 1.0 m solution would contain approximately 3.0 m solution of ions. In order to calculate colligative properties for electrolyte solutions, another factor (called the van't Hoff factor) is included in the equations.

In dilute solutions, the van't Hoff factor is equal to the number of ions that can be formed from each molecule. In concentrations solutions, sometimes the molecules do not all dissociate 100% and so the van't Hoff factor will be slightly less than the number of ions that can be formed from each molecules. Sometimes the values are referred as the theoretical and actual van't Hoff factors. In the absence of any information indicating the van't Hoff factor is less that the number of ions that can be formed from each molecule, assume full dissociation.

van't Hoff factor = i = number of moles of ions per mole of electrolyte

$i =$ actual number of particles in solution after dissociation

Example: What is the van't Hoff factor, i for each of the following solutions?

a. $Na_3PO_4 : i = 4$

b. $KOH : i = 2$

c. $Al(NO_3)_3 : i = 4$

d. $H_2SO_4 : i = 3$

For electrolyte solutions, the van't Hoff factor is included in the equations.

$$\Delta T_b = iK_bm$$
$$\Delta T_f = iK_fm$$
$$\pi = iMRT$$

Calculate the freezing point of a 0.100 m solution of CaS. K_f for water = $1.86°C/m$.

$$\Delta T_f = iK_fm = (2)(1.86°C/m)(0.100\ m) = 0.37°C$$
$$T_{fp} \text{ for the solution} = 0°C - 0.37°C = -0.37°C$$

Calculate the boiling point of 0.25 m solutions of $CaCl_2$. $K_b = 0.52°C/m$.

$$\Delta T_b = iK_bm = (3)(0.52°C/m)(0.25\ m) = 0.39°C$$
$$T_{bp} = 100.00°C + 0.39°C = 100.39°C$$

If the osmotic pressure of 0.010 M KI solution at $25°C$ is 0.465 atm, calculate the van't Hoff factor for KI at this concentration.

$$i = \frac{\pi}{MRT} = \frac{0.465\ atm}{(0.010\ mol/L)(0.0821\ L \cdot atm/mol \cdot K)(298\ K)} = 1.9$$

Determination of Molar Mass from Freezing Point Depression

Example: Ethylene glycol (EG) is a common automobile antifreeze. It is water soluble and a non-electrolyte. Calculate the molar mass of EG if 651 g of EG was dissolved in 2505 g of water, and the freezing point for the solution is $-7.79°C$. $K_f = 1.86°C/m$.

$$m = \frac{\Delta T_f}{K_f} = \frac{7.79°C}{1.86°C/m} = 4.19\ mol/kg$$
$$\text{mols} = (4.19\ mol/kg)(2.505\ kg) = 10.5\ mols$$
$$\text{molar mass} = \frac{grams}{mols} = \frac{651\ g}{10.5\ mols} = 62.0\ g/mol$$

Pacing the Lessons

Use the table below as a guide for the time required to teach the lessons of *The Solution Process*.

TABLE 21.1: 60 Minute Class Periods per Lesson

Lesson	Number of Class Periods
21.1 *What Are Solutions*	0.5
21.2 *Why Solutions Occur*	1.0
21.3 *Solution Terminology*	1.0
21.4 *Measuring Concentration*	2.0
21.5 *Solubility Graphs*	1.0
21.6 *Factors Affecting Solubility*	1.0
21.7 *Colligative Properties*	1.5
21.8 *Colloids*	0.5
21.9 *Separating Mixtures*	1.5

Managing Materials

The following items are needed to teach the strategies and activities described in the Teachers Edition of the FlexBook for *The Solution Process*.

TABLE 21.2: The Solution Process Materials List

Lesson	Strategy or Activity	Materials Needed
21.1	Engagement Activity	Vinegar, vegetable oil, beakers
21.2		
21.3		
21.4		
21.5		
21.6	Engagement Activity	20 oz. bottles of soda
21.7	Exploration Activity	Quart sized Ziploc bags, gallon sized Ziploc bags, milk, sugar, vanilla, salt, ice
21.8	Engagement Activity	Several solutions and colloids, beakers, black construction paper, and flashlight
21.9	Exploration Activity	Paper towels, straws, cups

Multimedia Resources

You may find these additional web based resources helpful when teaching *The Solution Process*:

- Virtual mixtures lab: http://www.harcourtschool.com/activity/mixture/mixture.html
- Cleaning water activity: http://acswebcontent.acs.org/games/clean_water.html
- Lesson on the factors affecting solubility: http://www.chem.lsu.edu/lucid/tutorials/solubility/Solubility.html

Possible Misconceptions

Identify: Students may think that the solubility of solid solute in a liquid solvent always increases with temperature.

Clarify: There are some exceptions to this trend.

Promote Understanding: Add 40 *g* of calcium acetate to 100 *mL* of water. Heat the solution until the calcium acetate precipitates out of the solution. Add ice to the solution and the solution will redissolve.

Making the FlexBook Flexible

An important advantage of the FlexBook is the ability it gives you, the teacher, to select the chapters and lessons that you think are most important for your own classes. You should also consult the standards correlation table that follows when selecting chapters and lessons to include in the FlexBook for your classes.

TABLE 21.3: Standard Addressed by the Lessons in The Solution Process

Lesson	California Standards	NSES Standards	AAAS Benchmarks
21.1			
21.2	6b		
21.3	6a		
21.4	6d		
21.5	6c		
21.6	6c		
21.7	6e		
21.8			
21.9	6f		

21.1 The Solution Process

Key Concepts

In this lesson, students explore the composition of solutions.

Lesson Objectives

- Define solutions.
- Describe the composition of homogeneous solutions.
- Describe the different types of solutions that are possible within the three states of matter.
- Identify homogeneous solutions of different types.

Lesson Vocabulary

solution A homogenous mixture; composition can vary; but composition is the same throughout once the solution is made.

Strategies to Engage

- Ask students to give examples of solutions. If students only give examples of solutions in solids, explain to them that solutions are possible with other states of matter as well.

Strategies to Explore

- Challenge students to fill in Table 21.1 with as many examples of actual solutions as they can think of. Award a prize to the student who can come up with the most (correct) examples.

Strategies to Extend and Evaluate

- Have each student record the four sentences in this section that most clearly represent the main ideas. Read key sentences in the text and have students raise their hands if they have recorded that sentence. Facilitate a discussion in which students defend their selections.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 21.1 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. What makes a solution homogeneous?
2. Which of the following are homogeneous? Explain.
   a. gasoline
   b. chocolate
   c. blood
   d. brass
3. Which of the following is a solution?
   a. milk
   b. blood
   c. gold
   d. air
   e. sugar
4. Which of the following is not a true solution?
   a. vinegar
   b. sand and water
   c. hard water, $CaCO_3(aq)$
   d. mercury alloy
5. Give an example of a homogeneous solution that is made from the following combinations
   a. a gas in a liquid
   b. a solid in a solid
   c. a solid in a liquid
   d. a gas in a gas
6. Jack is practicing some household chemistry. He takes 1 tsp of sugar and dissolves it in 250mL of water. He sees that the solution remains clear so continues his experiment by adding a second tsp of sugar. Stirring the solution makes this solution turn clear. After a few more attempts, Jack sees the solution turn murky then sugar crystals sinking to the bottom. What is Jack demonstrating?

21.2 Why Solutions Occur

Key Concepts

In this lesson students explore the relationship between molecular structure and solution formation.

Lesson Objectives

- Describe why solutions occur; the "like dissolves like" generalization.
- Determine if solutions will occur by studying the molecular structure.
- State the importance of water as the "universal solvent."

Lesson Vocabulary

intermolecular bonds Forces of attraction between molecules.

intramolecular bonds Forces of attraction between atoms in a molecule.

universal solvent A solvent able to dissolve practically anything (water).

Strategies to Engage

- Add a few drops of vinegar to a beaker of water, then add a few drops of vegetable oil to another beaker of water. Students should notice that the vinegar mixes with the water to form a solution while the vegetable oil does not. Facilitate a discussion with students in which they attempt to explain this occurrence. Explain to students that in this lesson, they will find out why this occurs.

Strategies to Explore

- Have students write a paragraph explaining Figure 21.1. Tell students to include the vocabulary terms in their explanations.
- Have students write a paragraph explaining Figure 21.1 in terms of the kinetic molecular theory.
- Explain to students that solutions in which water is the solvent are called aqueous solutions.

Strategies to Extend and Evaluate

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 21.2 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. What is the "like dissolves like" generalization and provide an example to illustrate your answer.
2. Why will $LiCl$ not dissolve in CCl_4?
3. Will acetic acid dissolve in water? Why?
4. What is the difference between intermolecular and intramolecular bonds?
5. In which compound will benzene (C_6H_6) dissolve?
   a. Carbon tetrachloride
   b. water
   c. vinegar
   d. none of the above
6. In which compound will sodium chloride dissolve?
   a. Carbon tetrachloride
   b. methanol
   c. vinegar
   d. none of the above
7. In which compound will ammonium phosphate dissolve?
   a. Carbon tetrachloride
   b. water
   c. methanol
   d. None of the above
8. Thomas is making a salad dressing for supper using balsamic vinegar and oil. He shakes and shakes the mixture but cannot seem to get the two to dissolve. Explain to Thomas why they will not dissolve.

21.3 Solution Terminology

Key Concepts

In this lesson students explore vocabulary associated with solutions.

Lesson Objectives

- Students will define solute, solvent, soluble, insoluble, miscible, immiscible, saturated, unsaturated, concentrated, and dilute.

Lesson Vocabulary

solute The substance in a solution present in the least amount.

solvent The substance in a solution present in the greatest amount.

soluble The ability to dissolve in solution.

insoluble The inability to dissolve in solution.

miscible Two liquids having the ability to be soluble in each other.

immiscible Two liquids not having the ability to be soluble in each other.

saturated A solution holding the maximum amount of solution in a given amount of solvent.

unsaturated A solution holding less than the maximum amount of solution in a given amount of solvent.

concentrated A solution where there is a large amount of solute in a given amount of solvent.

dilute A solution where there is a small amount of solute in a given amount of solvent.

Strategies to Engage

- Preview the lesson vocabulary to find out what your students already know about the concepts to be explored in this lesson. Have students define each vocabulary term. At the end of the lesson encourage students to go back and write the correct definition for each incorrect definition.

Strategies to Explore

- Have students research the Latin word *miscere* and write a paragraph relating it to the terms "*miscible*" and "*immiscible*".

Strategies to Extend and Evaluate

- Have students write questions derived from Bloom's Taxonomy. Instruct students to research Bloom's Taxonomy and write and answer one question from each of the six levels: knowledge, comprehension, application, analysis, synthesis, and evaluation.
- Ask each student to choose a set of lesson vocabulary terms such as solvent and solute, soluble and insoluble, miscible and immiscible, saturated and unsaturated, concentrated and dilute, and create a poster comparing and contrasting the two terms.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 21.3 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Distinguish between soluble, insoluble and miscible, immiscible. Use an example in your answer.
2. How can a solution that is concentrated be made more dilute and a dilute be made more concentrated?
3. Vinegar and water will mix together. Therefore two liquids are said to be:
    a. saturated
    b. miscible
    c. unsaturated
    d. immiscible
4. A solution is analyzed and found to contain 90 g of solute in 100 mL of solution. What can be concluded about this solution?
    a. The solution is concentrated.
    b. The concentration of the solution is 90 g/100 mL of water.
    c. The solution is saturated.
    d. The solution is holding the maximum amount of solute.
5. A solute is defined as:
    a. The substance in a solution present in the least amount.
    b. The substance in a solution that represents less than 50% of the solution.
    c. The substance that is dissolved in the solvent.

d. All of the above.
e. None of the above.

6. Match the following words with the examples that describe them.

a. solute - Adding only one can of water to a frozen concentrated juice mix will form this type of solution.
b. solvent - Adding eight cans of water to a frozen concentrated juice mix will form this type of solution.
c. soluble - Alcohol and water will have this property.
d. insoluble - Gasoline and water will have this property.
e. miscible - The water of a $NaOH(aq)$ solution.
f. immiscible - When salt is added to water it is said to have this property.
g. saturated - The copper(II) sulfate crystals in a solution of $CuSO_4(aq)$
h. unsaturated - The maximum amount of silver nitrate that can dissolve in 100 mL of water is 220 g. What term is given to this solution?
i. concentrated - If 220 g of $AgNO_3$ can dissolve in 100 g of water and only 50 g are added, what type solution is formed?
j. dilute - Adding calcium hydroxide to water forms a milky white precipitate. What term is given to calcium hydroxide?

7. Nisi is given two bottles of copper (II) sulfate solutions in her senior high chemistry lab. She is told that one bottle contains a saturated solution and the other one contains an unsaturated solution. What can Nisi do to identify the two solutions?
8. Can you have a solution that is saturated and dilute at the same time? Explain.

21.4 Measuring Concentration

Key Concepts

In this lesson students learn methods of expressing solution concentration.

Lesson Objectives

- Define molarity, mass percent, ppm, and molality.
- Calculate molarity, mass percent, ppm, and molality.
- Explain the importance of quantitative measurement in concentration.

Lesson Vocabulary

molarity A concentration unit measuring the moles of solute per liter of solution.

mass percent A concentration unit measuring the mass of solute per mass of solution. This unit is presented as a percent.

weight percent Another name for mass percent.

parts per million A concentration unit measuring the mass of solute per mass of solution multiplied by 1 million.

molality A concentration unit measuring the moles of solute per kilograms of solutions.

Strategies to Engage

- Have students read the review questions at the end of this section. This way, students will be familiar with the types of information that they will explore in this section.

Strategies to Explore

- Divide students into groups of three or four to work on problems in this lesson.
- Use the mathematical calculations in this lesson to reduce the reliance on language skills. As you go through each example problem, use them to explain the concepts explored in this lesson. DI **English Language Learners**

- Facilitate a discussion with students in which they compare and contrast the methods of expressing solution concentration explored in this lesson.

Strategies to Extend and Evaluate

- Have students use grid paper to make a crossword puzzle using the vocabulary terms. Ask students to exchange papers with a classmate and solve each other's puzzles.

Lesson Worksheets

Copy and distribute the lesson worksheets in the *CK-12 Chemistry Workbook* titled **Concentration by Percent Mass**, **Mole Fraction and Molality**, and **Molarity**. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 21.4 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Calculate the mass percent of silver when a silver/nickel solution is made with 34.5 g of silver and 72.3 g of nickel.
2. What would be the ppm of silver for the data presented in question 1?
3. Why is it a good idea to learn mass percent when molarity and molality are the most commonly used concentration measures?
4. Most times when news reports indicate the amount of lead or mercury found in foods, they use the concentration measures of *ppb* (parts per billion) or ppm (parts per million). Why use these over the others we have learned?
5. What is the molarity of a solution prepared by dissolving 2.5 g of $LiNO_3$ in sufficient water to make 60 mL of solution?
   a. 0.036 mol/L
   b. 0.041 mol/L
   c. 0.60 mol/L
   d. 0.060 mol/L
6. A solution is known to have a concentration of 325 ppm. What is the mass of the solute dissolved in 1.50 kg of solvent?
   a. 0.32 mg
   b. 0.49 mg
   c. 325 mg
   d. 488 mg
7. Calculate the molality of a solution of copper(II) sulfate where 11.25 g of the crystals has been dissolved in 325 g of water.
   a. 0.0346 m

b. 0.0705 m
c. 0.216 m
d. None of the above

8. What is the mass of magnesium chloride present in a 250 g solution found to be $21.4\% MgCl_2$?
   a. 21.4 g
   b. 53.5 g
   c. 196.5 g
   d. 250 g
9. What is the concentration of each of the following solutions in mol/L.
   a. 3.50 g of potassium chromate dissolved in 100 mL of water
   b. 50.0 g of magnesium nitrate dissolved in 250 mL of water.
10. Find the mass of aluminum nitrate required to produce 750 g of a 1.5 molal solution.
11. The Dead Sea contains approximately 332 grams of salt per kilogram of seawater. Assume this salt is all $NaCl$. Given that the density of the Dead Sea water is approximately 1.20 g/mL, calculate:
    a. the mass percent of $NaCl$.
    b. the mole fraction of $NaCl$.
    c. the molarity of $NaCl$.

21.5 Solubility Graphs

Key Concepts

In this lesson students explore the information provided by a solubility graph.

Lesson Objectives

- Students will read and report data from solubility graphs.
- Students will read and report saturation points from a solubility graph.

Lesson Vocabulary

solubility The amount of solute that will dissolve in a given amount of solvent at a particular temperature.

solubility graph A solubility graph is drawn to display the solubility at different temperatures. It is the mass of the $\frac{\text{solute}}{100\ g}$ of H_2O versus temperature in $°C$.

Strategies to Engage

Strategies to Explore

- Use the graphical nature of this lesson to reduce the reliance on language skills. As you go through each example problem, use the graphs to explain the concepts explored in this lesson. DI **English Language Learners**

Strategies to Extend and Evaluate

- Have students write a lesson to teach someone how to read a solubility graph. Tell students to include examples of each key concept.

Lesson Worksheets

There are no worksheets for this lesson.

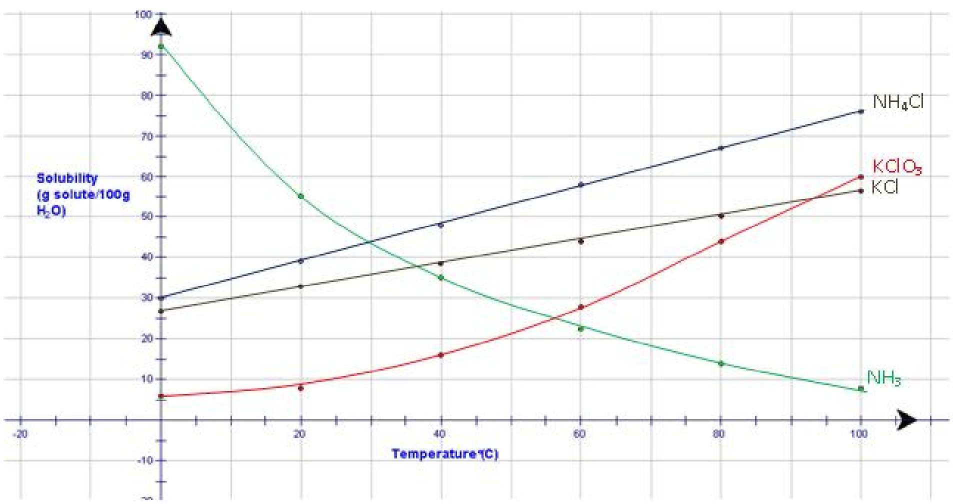
Review Questions

Have students answer the Lesson 21.5 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Using the graph below, determine:
    a. How much ammonia will dissolve at 30°*C*?
    b. What solid is more soluble at 50°*C*?
    c. What solid is least soluble at 60°*C*?
    d. At what temperature will 50 g of ammonium chloride dissolve in 100 g of water?



2. Why are solubility graphs useful?
3. Define solubility and solubility graph.
4. How many grams of *NaCl* are in 450 g of water at 30°*C* if the solubility is 39.8 g per 100 g of water?
    a. 8.84 g
    b. 39.8 g
    c. 100 g
    d. 179 g
5. How many moles of ammonium chloride are in 225 g of water at 40°*C* if the solubility is 45.8 g per 100 g of water?
    a. 0.86 mol
    b. 1.92 mol
    c. 20.3 mol
    d. 103 mol
6. How many moles of potassium chloride are in 500 g of water at 80°*C* if the solubility is 51.3 g per 100 g of water?
    a. 0.140 mol
    b. 0.688 mol
    c. 3.44 mol
    d. 10.3 mol

7. Plot the following data on a solubility graph and then answer the questions below.
   a. Which substance is the most soluble at 50°C?
   b. Which substance is the least soluble at 90°C?
   c. What is the solubility of NH_4ClO_4 at 30°C?
   d. How many grams of NH_4ClO_4 would dissolve in 250 mL at 30°C?
   e. At what temperature will 20 g potassium sulfate dissolve in 100 g of water?

TABLE 21.4:

Temp (°C)	g $NH_4Br/$100 g H_2O	g $NH_4ClO_4/$100 g H_2O	g $NaClO_3/$100 g H_2O
0	60.0	13.0	80.0
20	75.5	23.5	98.0
40	92.0	36.8	118.0
60	107.8	51.5	143.0
80	126.0	67.9	172.0
100	146.0	87.0	207.0

8. Plot the following data on a solubility graph and then answer the questions below.
   a. Which substance is the most soluble at 50°C?
   b. Which substance is the least soluble at 90°C?
   c. What is the solubility of $CuSO_4$ at 30°C?
   d. At what temperature will 20 g potassium sulfate dissolve in 100 g of water?

TABLE 21.5:

Temp (°C)	g $NaCl/$100 g H_2O g	$K_2SO_4/$100 g H_2O	g $CuSO_4/$100 g H_2O
0	35.7	7.4	14.3
20	36.0	11.1	20.7
40	36.5	14.8	28.7
60	37.3	18.2	40.0
80	38.1	21.4	56.0
100	39.2	24.1	80.0

21.6 Factors Affecting Solubility

Key Concepts

In this lesson students learn the factors that affect the solubility of solids and gases.

Lesson Objectives

- Describe the factors that affect solid solubility.
- Describe the factors that affect gas solubility.
- Describe how pressure can affect solubility.

Lesson Vocabulary

Henry's Law At a given temperature the solubility of a gas in a liquid is proportional to the pressure of that gas.

Strategies to Engage

- Have students observe 20oz bottles of warm and cold soda. Because the warm soda has less dissolved carbon dioxide, Students should notice that the warm soda has more space above the liquid and it may be a little wider. Open each bottle. Students should notice that the warm soda has a louder fizzing sound and more bubbles. Explain to students that by the end of this lesson they will be able to explain these occurrences.

Strategies to Explore

- Have groups of students design and conduct a scientific investigation on the effect of temperature and surface area on the solubility of sugar. Instruct students to come up with a list of necessary materials and equipment, and to write a step-by-step procedure. After the materials and procedure has been approved, have the groups conduct their investigations, and then write a lab report.

Strategies to Extend and Evaluate

- Challenge interested students to research the effects of thermal pollution on aquatic life and relate it to the concept of solubility. Students should be prepared to share their findings with the class.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 21.6 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. What are the factors that affect solubility?
2. What is Henry's Law?
3. Is it ever possible to have ionic solids decrease solubility with increasing temperature?
4. What factor would affect the solubility of sodium sulfate?
    a. i, ii, and iii
    b. i and ii
    c. i and iii
    d. ii and iii
        a. temperature
        b. pressure
        c. surface area
5. What factor would affect the solubility of methane?
    a. i, ii, and iii
    b. i and ii
    c. i and iii
    d. ii and iii
        a. temperature
        b. pressure
        c. surface area
6. If you crush a cube a sugar before putting it in your cup of coffee, how have you affected its solubility?
    a. Crushing it has really no affect on solubility because we have not heated it at all.
    b. Crushing it has increased the surface area so it speeds up the dissolving process but doesn't change maximum solubility.
    c. Crushing it has really no affect on solubility because we have not stirred it at all.
    d. Crushing it has increased the surface area so it increases the maximum solubility.
7. Why do people add chlorine to their swimming pools on a hot day?
8. Explain why crushed table salt at room temperature dissolves faster than rock salt.
9. Under which of the following sets of conditions would the solubility of $CO_2(g)$ be lowest? The pressure given is the pressure of $CO_2(g)$ above the solution.
    a. 5.0 atm and 75°C
    b. 1.0 atm and 75°C
    c. 5.0 atm and 25°C
    d. 1.0 atm and 25°C
    e. 3.0 atm and 25°C

10. An aqueous solution of KCl is heat from $15°C$ to $85°C$. Which of the following properties of the solution remain the same?
    a. i only
    b. iii only
    c. i and ii only
    d. ii and iii only
    e. i, ii, and iii
        a. molality
        b. molarity
        c. density

21.7 Colligative Properties

Key Concepts

In this lesson students learn the colligative properties of solutions and practice calculations involving them.

Lesson Objectives

- Describe vapor pressure lowering.
- Define boiling point elevation and freezing point depression.
- Describe what happens to the boiling points and freezing points, when a solute is added to a solvent.
- Describe the importance of the Van't Hoff factor.
- Calculate the boiling point elevation for electrolyte and non-electrolyte solutions.
- Calculate the freezing point depression for electrolyte and non-electrolyte solutions.

Lesson Vocabulary

boiling point elevation The difference in the boiling points of the pure solvent from the solution.

freezing point depression The difference in the freezing points of the solution from the pure solvent.

Van't Hoff factor The number of particles that the solute will dissociate into upon mixing with the solvent

Strategies to Engage

- Ask students why they think salt is used to melt ice on roads and sidewalks. Use this opportunity to gauge student understanding, clear up misconceptions, and generate curiosity for the concepts explored in this lesson.

Strategies to Explore

- Have less proficient readers make a main ideas/details chart as they read the lesson. Instruct them to divide a sheet of paper down the middle and record the main ideas on the left side and the details for each main idea on the right side. Have students save their chart for reviewing lesson content. DI **Less Proficient Readers**
- Divide students into groups of three or four to work on problems in this lesson.
- Use the mathematical calculations in this lesson to reduce the reliance on language skills. As you go through each example problem, use them to explain the concepts explored in this lesson. DI **English Language Learners**

- Have students work in pairs to make ice cream by following the following procedure: Add $\frac{1}{2}$ cup sugar, $\frac{1}{2}$ cup milk, and $\frac{1}{4}$ teaspoon vanilla to a quart size plastic bag and seal the bad securely. Put 2 cups of ice into a gallon size plastic bag and measure and record its temperature. Add $\frac{1}{2}$ cup of salt to the bag and measure and record the temperature again. Place the quart size plastic bag inside of the gallon bag and seal the gallon bag. Gently massage the bag for about 25 *minutes*. During the mixing process, facilitate a discussion with students about why the temperature of the ice/salt mixture was lower than the ice alone.

Strategies to Extend and Evaluate

- Have students bring in examples of applications of colligative properties in everyday life; such as adding salt to water in order to increase its boiling point while cooking. Students should be prepared to share their findings with the rest of the class.

Lesson Worksheets

Copy and distribute the lesson worksheets in the *CK-12 Chemistry Workbook* titled **Solution Vapor Pressure Lowering**, **BP Elevation and MP Depression**, and **Dilution**. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 21.7 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. What must be measured in order to determine the freezing point depression?
2. From a list of solutions with similar molalities, how could you quickly determine which would have the highest boiling point?
3. Why would table salt not be a good solution to use when deicing a plane?
4. When a solute is added to a solution:
    a. i and iii are true
    b. i and iv are true
    c. ii and iii are true
    d. ii and iv are true
        a. the boiling point increases.
        b. the boiling point decreases.
        c. the freezing point increases.
        d. the freezing point decreases.
5. If 25.0 g of sucrose $(C_{12}H_{22}O_{11})$ is added to 500. g of water, the boiling point is increased by what amount? $(K_b\ (\text{ water}) = 0.52^\circ\ \text{C/m})$
    a. 0.076°
    b. 0.025°
    c. 26°

d. None of these

6. The solubility of seawater (an aqueous solution of $NaCl$) is approximately 0.50 m. Calculate the freezing point of seawater. $(K_f \text{ (water)} = 1.86^\circ \text{ C/m})$

    a. -0.93°
    b. 0.93°
    c. 1.86°
    d. -1.86°

7. Determine which of the following solutions would have the lowest freezing point.

    a. 15 g of ammonium nitrate in 100. g of water.
    b. 50. g of glucose in 100. g of water.
    c. 35 g of calcium chloride in 150. g of water.

8. A 135.0 g sample of an unknown nonelectrolyte compound is dissolved in 725 g of water. The boiling point of the resulting solution was found to be $106.02^\circ C$. What is the molecular weight of the unknown compound?
9. What is the van't Hoff factor for each of the following:

    a. $MgCl_2$
    b. Ammonium sulfate
    c. CH_3OH
    d. Potassium chloride
    e. KCH_3COO

10. Calcium chloride is known to melt ice faster than sodium chloride but is not used on roads because the salt itself attracts water. If 15 g of $CaCl_2$ was added to 250 g of water, what would be the effect on the freezing point of the solution? $(K_f(\text{ water}) = -1.86^\circ \text{ C/m})$

21.8 Colloids

Key Concepts

In this lesson students explore the similarities and differences among solutions, colloids, and suspensions.

Lesson Objectives

- Define colloids and suspensions.
- Compare solutions, colloids, and suspensions.
- Characterize solutions as suspensions, colloids, or solutions.
- Name some common examples of colloids.

Lesson Vocabulary

colloid Mixtures where the size of the particles is between 1×10^3 *pm* and 1×10^6 *pm* (i.e., milk).

suspension Mixtures where the particles settles to the bottom of the container and can be separated by filtration.

Tyndall Effect Involves shining a light through the mixture, if the light scatters, the mixture is a colloid.

Strategies to Engage

- Place several solutions (such as salt/water and soda) and colloids (such as milk and cornstarch/water), into separate beakers. Label each beaker with the name of the material it contains. Make a cone from black construction paper and tape it over the lens of a flashlight. Turn off the lights in the room and shine the narrow beam of light at each of the beakers. The beam of light will be visible in the colloids, but will not be visible in the solutions. Tell the students that by the end of this lesson they will be able to explain these occurrences.

Strategies to Explore

- Point out to students that the main difference between solutions, colloids, and suspensions is the size of the particles. Solutions have the smallest particle size, followed by colloids. Suspensions have the largest particle size.

Strategies to Extend and Evaluate

- Have students organize the information explored in this lesson into a table that summarizes the properties of solutions, colloids, and suspensions. Ask students to include examples of each and other information such as particle size and Tyndall effect.
- Have students create a poster that includes examples of edible solutions, colloids, and suspensions.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 21.8 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Distinguish between a solution, a colloid, and a suspension.
2. What is one true way to tell you have a colloid solution?
3. Why do you think there is no example of a gas - gas colloid?
4. Which is an example of a colloid?
   a. air
   b. brass
   c. milk
   d. none of these
5. Which is not an example of a colloid?
   a. human body
   b. mayonnaise
   c. mustard
   d. cloud
6. The biggest difference between a colloid and a suspension is that:
   a. In colloids, the solute is permanently dissolved in the solvent.
   b. In colloids the particles eventually settle to the bottom.
   c. In suspensions the particles eventually settle to the bottom.
   d. None of these are correct
7. Karen was working in the lab with an unknown solution. She noticed that there was no precipitate in the bottom of the beaker even after it had been on the lab bench for several days. She tested it with a light and saw that light scattered as it passed through the solution. Karen concluded that the liquid was what type of a mixture?
   a. colloid
   b. suspension
   c. homogeneous
   d. heterogeneous

8. What are two good common examples of colloids?

21.9 Separating Mixtures

Key Concepts

In this lesson students explore methods used to separate mixtures.

Lesson Objectives

- The students will describe differences between the physical properties of pure substances and solutions.
- The students will list and describe methods of separation for mixtures.
- The students will explain the principles involved in chromatographic separation.
- The students will identify the mobile and stationary phases in a chromatography design.
- Given appropriate data, the students will calculate Rf values.

Lesson Vocabulary

distillation The evaporation and subsequent collection of a liquid by condensation as a means of purification.

fractional distillation This is a special type of distillation used to separate a mixture of liquids, using their differences in boiling points.

chromatography Any of various techniques for the separation of complex mixtures that rely on the differential affinities of substances for a mobile solvent and a stationary medium through which they pass.

Strategies to Engage

- Have students read the review questions at the end of this section. This way, students will be familiar with the types of information that they will explore in this section.

Strategies to Explore

- Students can perform this simple chromatography experiment using a paper towel, a black washable marker, a straw, and a cup. Use the marker to draw a circle on the paper towel. Use a straw to add drops of water to the center of the circle. Students should be able to see the individual colors in the ink. Encourage students to perform the experiment again using different materials such as a coffee filter instead of a paper towel, alcohol instead of water, and drink mix instead of a marker.
- Have students create a chart that summarizes each of the separation methods explored in this lesson.

Strategies to Extend and Evaluate

- As a class, create a concept map of the information explored in this chapter.
- Have students complete the lab ***Separation of a Mixture***. This lab is located in the Supplemental Lab Book.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 21.9 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. In a paper chromatography experiment to separate the various pigments in chlorophyll, a mixture of water and ethanol was used as the solvent. What is the stationary phase in this separation?
2. Do you think that paper chromatography or TLC would be useful for separating a very large quantity of a mixture? Explain why or why not.
3. If the mobile phase in a chromatographic experiment moved 15.0 cm and one of the compounds in the mixture moved 12.7 cm, what is the R_f value for this compound?
4. If the stationary phase in a paper chromatography experiment was very polar and the solvent was moderately polar, would the polar components in the mixture be closer to the bottom of the paper or toward the top of the paper?

21.10 Enrichment

Extra Readings

Solutions

We are all familiar with the phenomenon of a hard crystalline solid, like table salt, when placed in water, apparently disappearing quite quickly. The crystalline structure breaks up and the particles enter into the water. Why does this process occur?

More questions arise when we think about dissolving and solutions. Table salt, for example, dissolves in water, but it will not dissolve in benzene. Camphor, on the other hand, dissolves easily in benzene, but not in water. While other substances like diamonds or graphite will not dissolve in any liquid. What controls whether a solid dissolves, and in what solvent it will dissolve?

Ion-Ion Attraction vs. Ion-Solvent Attraction

Consider the example of salt dissolving in water. Recall that table salt (sodium chloride) has a simple crystal structure in which positive sodium ions and negative chloride ions are organized in a crystal lattice. The electrical interactions between the positive and negative ions causes them to be strongly held at their locations in the crystal. To break up the crystal requires a large amount of energy, or else the attraction between the ions must be replaced by some other equal or greater attraction. This is the key to understanding what happens when the ions dissolve in water. The attraction between the ions in the solid is replaced by an attraction between the ions and the water molecules (or other solvent molecules). Water is a polar liquid. The oxygen end of the molecule has a partial negative charge, while the hydrogen end has a partial positive charge. When the sodium ion enters the liquid water, the water molecules cluster around it so that the partially negative ends of the water molecules are next to the positive sodium ions. Similarly the water molecules cluster around the chloride ions so that the partially positive ends of water molecules are directed toward the negative charge of the chloride ions. It is these ion-water attractions in the solution that replace the ion-ion attractions in the solid. The ions can break away from their oppositely charged neighbors in the crystal because they have found equal or stronger attractions in the solution. High solubility requires that the attraction between the atoms, ions, or molecules in the dissolving solid be replaced by equivalent or greater attractions between these particles and the molecules of the solvent. In many cases, it still requires an input of energy for a solid to dissolve in a solvent, but if the requirement is small enough, its effect can be outweighed by that of the increased disorder of the solution. The process is then driven by the increased entropy of the solution.

Like Dissolves Like

Solids like salt, which consist of ions, dissolve in polar solvents like water in which the solvent molecules have dipoles, because the electrical attractions between the ions and the solvent replace those between ions in the solid. In crystals made of non-polar molecules like camphor, the forces are different. The molecules are held in the crystal by weak London dispersion force attractions. Similar forces exist between the molecules in a solvent like benzene. So again, the interactions between the molecules in the solid can be replaced by those between the solute molecule and solvent molecules. Hence, camphor dissolves in benzene.

Much of what we have described so far relates to water as a solvent. Water is the most widespread liquid on the surface of earth. As we have seen, water is also an excellent solvent for polar solids. Polar solids include not only those that are made of ions like sodium chloride, but also those that are composed of polar molecules, like glucose and alcohol. In some cases, the attractions between solute molecules and the solvent water are even greater due to

hydrogen bonding.

Molecules or parts of molecules can be classified as hydrophilic (water loving) or hydrophobic (water hating) depending on whether they contain polar groups. An important group of molecules of this type are soaps and detergents, which have both hydrophilic and hydrophobic ends, leading to a range of useful and remarkable properties.

Solid Solutions

We do not normally think of solids like copper or silicon as being able to dissolve because there are no common liquids in which these solids will dissolve. Sometimes, the reaction between metals and strong acids are referred to as "dissolving", but that involves a chemical reaction. In order for a solid to dissolve, the interaction between the atoms or molecules in the solid must be replaced by comparable ones in the solution. There are no substances which are liquids at normal temperatures, in which the atoms or molecules attract metal atoms strongly enough to dissolve them (assuming no chemical reaction).

It is possible, however, for these solids to dissolve in other solids forming solid solutions. Copper will dissolve in zinc to form an alloy (a solution of one metal in another) known as brass. Like most alloys, brass is crystalline, that is, it has a regular arrangement of metal atom locations with some of the locations occupied by copper and some by zinc atoms. Alloy formation is very common; other examples are pewter (tin and zinc) and bronze (iron and copper). Dissolving a small amount of one metal in another can also have significant effects on physical and chemical properties. Stainless steel is essentially iron into which a small amount of chromium is dissolved. Stainless steel is significantly different from iron in terms of the rate at which it corrodes. Alloying iron with copper in bronze results in a much tougher, less brittle material. Solid solutions are widespread. Silicon, as we have seen, will not dissolve in any common liquid; but it will dissolve in germanium (a solid with the same crystal structure) in much the same way that copper dissolves in zinc. These solids further illustrate the point that dissolving small amounts of one solid substance in another is a vitally important way of altering the properties of materials. One that is used on an enormous scale in contemporary technology. The classic example is the semiconductor silicon. Dissolving tiny amounts (less than one part per million) of phosphorus in silicon has a significant effect on its ability to conduct electricity (making the material that is known as an 'n-type' semiconductor). Similar amounts of arsenic dissolved in silicon have equally large effects, but result in different electrical characteristics (the material becomes a 'p-type' semiconductor). Putting the two types of material together creates the famous p/n junction which allows electricity to flow only in one direction, a vital feature of some electrical circuits. Silicon with tiny quantities of deliberately introduced impurities is therefore the material basis of the technology on which the modern electronics revolution is based.

21.11 References

1. Richard Parsons. . CCBYSA
2. Richard Parsons. . CCBYSA
3. Richard Parsons. . CCBYSA

CHAPTER 22 Ions in Solution TE

Chapter Outline

Ions in Solutions

Outline

This chapter, *Ions in Solution* consists of three lessons that cover dissociation, electrolytes and non-electrolytes, reactions between ions in solution, and ionic and net-ionic equations.

- **Lesson 22.1 Ionic Solutions**
- **Lesson 22.2 Covalent Compounds in Solution**
- **Lesson 22.3 Reactions Between Ions in Solutions**

Overview

In these lessons, students will explore:

- What happens when ionic solids dissolve in water.
- What happens when covalent compounds dissolve in water.
- Solubility and the reactions between ions in solutions.

Science Background Information

This background information is provided for teachers who are just beginning to teach in this subject area.

Modern Armor from a Solution

Stephanie Kwolek graduated from the women's college of a much larger all-men's university. (Today, the two colleges form the co-ed Carnegie Mellon University.) With her degree in chemistry, Kwolek accepted a job at the DuPont Chemical Company. DuPont had been highly successful with its development of Nylon®, Dacron®, and other synthetic fibers and in the early 1960's, Kwolek was working on the development of new fibers. The process for developing new fibers at that time was to combine substances to make a polymer, melt the polymer into a liquid, and then spin the liquid in a machine called a "spineret". The liquid would squirt out through holes in the spineret and solidify into fibers.

Kwolek was directed to search specifically for high-performance fibers that were very stiff and strong and could be used to reinforce tires in place of steel wires. Lightweight fibers that were stiff and strong and resistant to high temperatures would have many profitable applications.

One day, Kwolek was experimenting with a polymer that was extremely difficult to melt and therefore couldn't be "spun" in the spineret. Kwolek decided to find a solvent that would dissolve the polymer and get it into liquid form in that manner, rather than melt it. After many tries, she eventually found a solvent that would dissolve the polymer. She had difficulty convincing the scientist who ran the spineret to "spin" her solution, because he felt that the solution would plug the holes in his machine. After several days, Kwolek finally convinced him to spin her solution. The solution spun beautifully and produced fibers that were very strong and very stiff. Kwolek baked the fibers and after baking, they were even stronger and stiffer.

Kwolek had made two discoveries. The solution she had produced was a new type of substance called liquid crystal solutions, and the fiber she had produced was a new kind of fiber called an aramid fiber. Para-aramid fibers go by the commercial name Kevlar®.

Today, aramid fibers are used to make bullet-resistant vests, boat hulls, coats, cut-resistant gloves, fiber-optic cables, firefighters' suits, fuel hoses, helmets, parts of airplanes, radial tires, special ropes, pieces of spacecraft, tennis rackets, canoes, and skis. Aramid fibers are stronger and lighter than steel. A vest made of seven layers of aramid fiber weighs about 2.5 *pounds* and can deflect both knife blades and bullets fired from a distance of 10 $feet$.

Pacing the Lessons

Use the table below as a guide for the time required to teach the lessons of *Ions in Solution*.

TABLE 22.1: 60 Minute Class Periods per Lesson

Lesson	Number of Class Periods
21.1 *Ionic Solutions*	0.5
21.2 *Covalent Compounds in Solution*	0.5
21.3 *Reactions Between Ions in Solution*	1.5

Managing Materials

The following items are needed to teach the strategies and activities described in the Teachers Edition of the FlexBook for *Ions in Solution*.

TABLE 22.2: Ions in Solution Materials List

Lesson	Strategy or Activity	Materials Needed
22.1		
22.2		
22.3		

Multimedia Resources

Possible Misconceptions

Making the FlexBook Flexible

An important advantage of the FlexBook is the ability it gives you, the teacher, to select the chapters and lessons that you think are most important for your own classes. You should also consult the standards correlation table that follows when selecting chapters and lessons to include in the FlexBook for your classes.

TABLE 22.3: Standard Addressed by the Lessons in Ions in Solution

Lesson	California Standards	NSES Standards	AAAS Benchmarks
22.1			
22.2			
22.3			

22.1 Ions in Solution

Key Concepts

In this lesson students explore what happens when ionic solids dissolve in water.

Lesson Objectives

- Describe electrostatic attraction.
- Explain how ionic solids attract water molecules when they dissolve in water.
- Explain the difference between physical changes and chemical changes.
- Define electrolyte solutions and be able to identify electrolytes.

Lesson Vocabulary

electrostatic attraction When solids form from a metal atom donating an electron (thus forming a positive cation) to a non-metal (thus forming a negative anion) the two ions in the solid are held together by the attraction of oppositely charged particles.

chemical changes Changes that occur with the chemical bonding where a new substance is formed.

physical changes Changes that occur in the physical structure but do not occur at the molecular level.

electrolyte solutions Solutions that contain ions that are able to conduct electricity.

Strategies to Engage

- Ask students to describe what happens when an ionic compound dissolves in water. Use this opportunity to gauge student understanding, address misconceptions, and generate curiosity for the concepts explored in this lesson.

Strategies to Explore

- Use Figure 22.4 to explain as many concepts as possible. Relate concepts such as electrostatic attraction and dissociation to Figure 22.4. DI **English Language Learners**

- Have less proficient readers make a main ideas/details chart as they read the lesson. Instruct them to divide a sheet of paper down the middle and record the main ideas on the left side and the details for each main idea on the right side. Have students save their chart for reviewing lesson content. DI **Less Proficient Readers**
- Perform the ***Conductivity of Solutions*** *demonstration. This demonstration is located in the Supplemental Lab Book.*

Strategies to Extend and Evaluate

- Challenge students to come up with their own question for each of the sections in this lesson. Students may then exchange papers with a classmate to have them answer each others questions.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 22.1 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Write the reactions for the dissolving of the following.
   a. $NaOH_{(s)}$
   b. $LiOH_{(s)}$
   c. $C_5H_{10}O_{4(s)}$
   d. $NH_4Cl_{(s)}$
   e. $MgCl_{2(s)}$
2. Which of the following represent physical changes? Explain.
   a. explosion of TNT
   b. dissolving KCl
   c. sharpening a pencil
   d. souring milk
3. Which compound contains electrostatic forces?
   a. natural gas
   b. table salt
   c. air
   d. sugar
4. Which of the following is a physical change?
   a. rotting wood
   b. rising of bread dough
   c. rusting iron
   d. molding cheese

5. Which of the following is not a physical change?
   a. melting iron
   b. pumping gas
   c. reaction of chlorine with sodium
   d. reaction of magnesium chloride with water
6. Which compound is considered to be an electrolyte when dissolved in water?
   a. HNO_3
   b. $C_{12}H_{22}O_{11}$
   c. N_2O
   d. CH_4
7. Which compound is not considered to be an electrolyte?
   a. $AgCl$
   b. $PbSO_4$
   c. C_2H_6
   d. $HClO_3$
8. Janet is given three solutions. She is to determine if the solutions are electrolytes or not but is not told what the solutions are. She makes the following observations. What can you conclude from her observations and what help can you offer Janet to determine if the solutions are indeed electrolytes? Solution 1: Clear Solution 2: Blue but transparent Solution 3: Clear

22.2 Covalent Compounds in Solution

Key Concepts

In this lesson students explore what happens when covalent compounds dissolve in water.

Lesson Objectives

- Describe intermolecular bonds.
- Explain why molecules stay together when dissolving in solvents.
- Define and explain non-electrolytes.

Lesson Vocabulary

intermolecular bonding The bonding that occurs between molecules.

non-electrolytes Solutions that do not conduct electricity.

Strategies to Engage

- Introduce lesson concepts by asking students to recall what they know about the similarities and differences between ionic and covalent compounds. Guide them in focusing their prior knowledge.

Strategies to Explore

- Ask students to write a one-paragraph summary of this lesson that demonstrate mastery of each lesson objective.

Strategies to Extend and Evaluate

- Have each student record the four sentences in this section that most clearly represent the main ideas. Read key sentences in the text and have students raise their hands if they have recorded that sentence. Facilitate a discussion in which students defend their selections.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 22.2 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Describe the intermolecular bonding that would occur between glucose, $C_6H_{12}O_6$, and water.
2. Define non-electrolyte and give at least one example.
3. How can you tell by looking at a formula that it is most likely a covalent compound? What does this tell you about the bonding?
4. Describe how you could tell the difference between an electrolyte and a non-electrolyte solution.
5. Looking at the periodic table, which pair of elements will form a compound that is covalent?
   a. Ca and Br
   b. Fe and O
   c. Si and F
   d. Co and Cl
6. Which of the following compounds will conduct the least amount of electricity if dissolved in water?
   a. KNO_3
   b. $BaCl_2$
   c. CsF
   d. CO_2
7. Steve is given five solutions in the lab to identify. He performs a conductivity test, a solubility (in water) test, crudely measures the hardness of each substance, and determines the melting point using a melting point apparatus. Some of the melting points, the teacher tells him are too high or low to measure using the laboratory melting point apparatus so she gives him the melting point. For the liquids, he determined the boiling points. He gathers all of his data and puts it into a table. His teacher gives him the names of the five solutions to match his five unknowns to. Can you help Steve match the properties of the unknowns (from the table below) to the solution names (found under the table)?

TABLE 22.4:

Unknown Substance	Conductivity	Solubility (in water)	Hardness	Melting Point (°C)	Boiling Point (°C)
1	no (aq)	soluble	semi- brittle	164	
2	yes (aq)	soluble	NA(liquid)		100
3	yes (aq)	soluble	brittle	≈ 800	
4	no (s)	insoluble	soft	82	
5	yes (s)	soluble	NA(liquid)		118

List of Unknown names:

Sodium chloride

Naphthalene

Sucrose

Hydrochloric acid (dilute)

Acetic acid

8. Predict the type of bonding that will form between the elements sulfur and bromine. Will this molecule conduct electricity in water solution?

22.3 Reactions Between Ions in Solutions

Key Concepts

In this lesson students explore solubility and the reactions between ions in solutions.

Lesson Objectives

- Use the solubility chart and/or solubility rules to determine if substances are soluble in water.
- Use the solubility chart and/or the solubility rules to determine if precipitates will form.
- Write molecular, ionic, and net ionic equations.
- Identify spectator ions in ionic equations.

Lesson Vocabulary

solubility chart A grid showing the possible combinations of cations and anions and their solubilities in water.

solubility rules A list of rules dictating which combinations of cations and anions will be soluble or insoluble in water.

formula equation A chemical equation written such that the aqueous solutions are written in formula form.

total ionic equation A chemical equation written such that the actual free ions are shown for each species in aqueous form.

net ionic equation The overall reaction that results when spectator ions are removed from the ionic equation.

spectator ions The ions in the total ionic equation that appear in the same form on both sides of the equation indicating they do not participate in the overall reaction.

Strategies to Engage

Strategies to Explore

- Have students write step-by-step instructions for writing net ionic equations and use these steps to complete the practice problems in this lesson.

Strategies to Extend and Evaluate

- Have students complete the lab ***Qualitative Ion Testing***. This lab is located in the Supplemental Lab Book.

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Reactions Between Ions in Solution**. Ask the students to complete the worksheet alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 22.3 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. What is more valuable to use for determining solubility: a solubility chart or a set of solubility rules?
2. If you were told to visualize $Cu(NO_3)_2(aq)$, what might this mean to you?
3. Use the solubility rules to determine the following solubilities in water. **Solubility Rules**
   a. All group 1 metals and ammonium compounds are soluble.
   b. All nitrates, chlorates, and bicarbonates are soluble.
   c. Halides are soluble except for Ag^+, $Hg_2{}^{2+}$, and Pb^{2+}
   d. Sulfates are soluble except for Ag^+, Ba^{2+}, Ca^{2+}, $Hg_2{}^{2+}$, Sr^{2+}, and Pb^{2+}
   e. Carbonates, chromates, phosphates, and sulfides are insoluble except those from rule #1.
   f. Hydroxides are insoluble except for those in rule #1, and Ba^{2+}.

   Which of the following compounds is soluble in water?

   a. $PbCl_2$
   b. Hg_2Cl_2
   c. $(NH_4)_2SO_4$
   d. $MgCO_3$
   e. $AgNO_3$
   f. $MgCl_2$
   g. KOH
   h. $PbSO_4$
4. When only the ions that produce a precipitate are shown for a chemical equation, what type of reaction exists?
   a. spectator equation
   b. molecular equation
   c. ionic equation
   d. net ionic equation
5. If you wanted to separate a solution of $Ag^{3+}{}_{(aq)}$ from a solution of $Hg_2{}^{2+}{}_{(aq)}$, which of the following would be the best possible reaction?
   a. add HBr
   b. add HNO_3
   c. add $NaOH$

d. none of the above

6. If you wanted to separate a solution of $Al^{+}{}_{(aq)}$ from a solution of $Fe^{2+}{}_{(aq)}$, which of the following would be the best possible reaction?
   a. add HCl
   b. add $NaOH$
   c. add H_2S
   d. none of the above
7. Complete the following reactions:
   a. $Na_2S(aq) + ZnCl_2(aq) \rightarrow$
   b. $(NH_4)_2CO_3(aq) + CaCl_2(aq) \rightarrow$
8. Write the ionic equations for the balanced molecular equations from question 5.
9. Write the net ionic equations for the ionic equations from question 6.
10. Identify the spectator ions for the ionic equations from question 6.

CHAPTER 23 Chemical Kinetics TE

Chapter Outline

Unit 8 Chemical Kinetics and Equilibrium

Outline

This unit, Chemical Kinetics and Equilibrium, includes two chapters that explore reaction rates and equilibrium.

- **Chapter 18 Chemical Kinetics**
- **Chapter 19 Chemical Equilibrium**

Overview

Chapter 18: Chemical Kinetics

This chapter covers reaction rates and the factors that affect reaction rates.

Chapter 19: Chemical Equilibrium

This chapter covers relationships between forward and reverse reaction rates, the concept of chemical equilibrium, the mathematics of the equilibrium constant, Le Chatelier's principle, and solubility product constant calculations.

Chemical Kinetics

Outline

This chapter *Chemical Kinetics* consists of five lessons that cover reaction rates and the factors that affect reaction rates.

- **Lesson 18.1 Rate of Reactions**
- **Lesson 18.2 Collision Theory**
- **Lesson 18.3 Potential Energy Diagrams**

- **Lesson 18.4 Factors That Affect Reaction Rates**
- **Lesson 18.5 Reaction Mechanism**

Overview

In these lessons, students will explore:

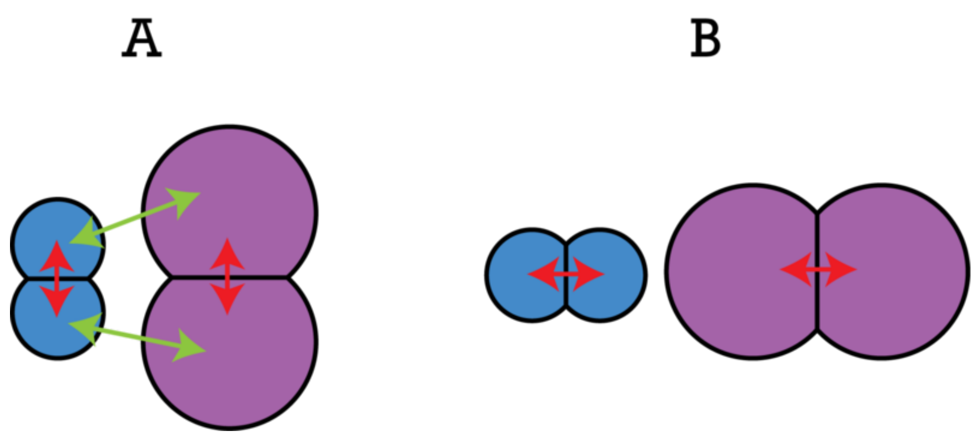
- the rates of chemical reactions.
- reactions rates in terms of collisions between reacting particles.
- potential energy diagrams for endothermic and exothermic reactions.
- the effect of temperature, surface area, concentration, and catalysts on the rate of a chemical reaction.
- multi-step processes as well as the individual reactions in a multi-step process.

Science Background Information

This information is provided for teachers who are just beginning to instruct in this subject area.

Chemical Kinetics

Chemical Kinetics is the study of the mechanisms and rates of chemical reactions. In order for a reaction to occur, a collision between reacting particles must occur. Assuming the reactant is a molecule, the atoms in the molecule are already bonded to at least one other atom. If the atom is to form new bonds during the reaction, the old bonds must first be broken. In some cases, the collision to break the old bonds must also have proper orientation. For example, consider the reaction between H_2 and I_2 shown below.



During the collision between H_2 and I_2, the $H-H$ bond and the $I-I$ bond (indicated by red arrows) must be broken and $H-I$ bonds (indicated by green arrows) must be formed. In collision A, the side-to-side collision of the molecules would require the least amount of energy to break the old bonds and the atoms would be in convenient position to form HI molecules. In collision B, however, the end-to-end collision would appear to push the atoms together rather than break them apart, and the H and I atoms in the outside positions are not in convenient position to bond. If collision B were to result in a reaction, a great deal more energy would be required than that for collision A.

Reaction rates are affected by the concentrations of the reacting species, the temperature of the reaction, and whether or not a catalyst is present. Reaction rates (at constant temperature) can be expressed in a mathematical

expression relating the rate of a reaction to the concentrations of the reactants. This rate law can be determined from experimental data.

Here is an example of an overall chemical reaction and the rate law for that reaction.

$$2\,NO_{(g)} + O_{2(g)} \rightarrow 2\,NO_{2(g)}$$
$$\text{Rate} = k[NO]^2[O_2]$$

The rate expresses the rate of production of NO_2 in moles/liter/sec or M/s, and is proportional to the concentrations of the reactants where k is a proportionality constant called the reaction constant. The exponent associated with each reactant is referred to as the order of the reaction with respect to that reactant. In this case, the reaction is 2^{nd} order with respect to NO and 1^{st} order with respect to O_2. The overall order of the reaction is the sum of partial orders with respect to each reactant. In this case, the overall order of the reaction is 3^{rd} order.

The reaction mechanism is the series of collisions that describe the steps involved in the reaction. Consider the following reaction.

$$CO + NO_2 \rightarrow CO_2 + NO$$

In this reaction, it has been experimentally determined that this reaction takes place according to the rate law $R = k[NO_2]^2$. Therefore, a possible mechanism by which this reaction takes place is:

$$2\,NO_2 \rightarrow NO_3 + NO \qquad \text{(slow)}$$
$$NO_3 + CO \rightarrow NO_2 + CO_2 \qquad \text{(fast)}$$

The first collision in this reaction occurs between two NO_2 molecules and produces an NO_3 molecule and an NO molecule. The second collision in the mechanism occurs between the NO_3 molecule produced in step 1 and a CO molecule producing an NO_2 and a CO_2. When all the steps in the reaction mechanism are added, the NO_3 (on both sides) cancel and NO_3 does not appear in the net reaction. The NO_2 in the product cancels one of the NO_2 molecules in the reactant and the net reaction is $CO + NO_2 \rightarrow CO_2 + NO$. The overall reaction rate for this reaction (and all reactions) will be exactly the same as the reaction rate for the slowest step. The rate law for the slowest step is $R = k[NO_2]^2$ and therefore, the rate law for the net reaction is the same, even though two NO_2 molecules do not appear in the reactants for the net reaction. That is why the rate law for the net reaction must be determined experimentally.

Some chemical reactions may occur with a single collision between reactant particles. The possibility of a single collision reaction is limited to reactions involving two particles or in some cases, three particles. The probability of three particles arriving at the same point at the same time for a single three-particle collision is low. Collisions involving more than three particles essentially never occur.

Suppose the reaction between carbon and oxygen to yield carbon dioxide occurred with a single collision between a carbon atom and an oxygen molecule.

$$C + O_2 \rightarrow CO_2$$

In such a case, the reaction mechanism and the net reaction are the same reaction. The net reaction represents the reaction mechanism and the slowest step in the reaction mechanism. Therefore, for this very simple reaction, the rate law may be written by looking at the net reaction; $\text{Rate} = k[C][O_2]$.

If a three particle reaction occurred in a single collision, the rate law could also be written from the net reaction.

$$2\,NO_{(g)} + O_{2(g)} \rightarrow 2\,NO_{2(g)}$$
$$\text{Rate} = k[NO]^2[O_2]$$

This reaction and the rate law could also be written in the following manner,

$$NO_{(g)} + NO_{(g)} + O_{2(g)} \rightarrow 2\,NO_{2(g)}$$
$$\text{Rate} = k[NO][NO][O_2]$$

and that's why the coefficients of the reactants become exponents in the rate law.

The great majority of reactions that involve more than two particles as reactants occur by a series of collisions (reaction mechanism) and for these reactions, the rate law must be determined experimentally.

Consider the following set of experimental data from which the rate law may be determined for the reaction between NO and O_2.

TABLE 23.1: Rate Data Table for the Reaction Between

Trial	Initial $[NO]$	Initial $[O_2]$	Experimentally Determined Rate
1	$0.10\,M$	$0.10\,M$	$1.2 \times 10^{-8}\,M/s$
2	$0.10\,M$	$0.20\,M$	$2.4 \times 10^{-8}\,M/s$
3	$0.30\,M$	$0.10\,M$	$1.08 \times 10^{-7}\,M/s$

We pick two trials in which one of the reactant concentrations is held constant and the other reactant concentration changes. To begin, we choose trials 1 and 2 in which the concentration of NO is constant and the concentration of O_2 changes. We can determine the order of the reaction with respect to O_2 with the following mathematics.

(multiple of O_2 concentration)^x = (multiple of rate), where the exponent, x, is the order of the reaction with respect to O_2.

The concentration of O_2 has been doubled and the rate has been doubled, so

$2^x = 2$ and therefore, $x = 1$. The order of the reaction with respect to oxygen is 1.

We now choose two trials in which the O_2 concentration is held constant and the concentration of NO varies. In trials 1 and 3, the concentration of NO has been tripled and the rate has been increased by a factor of 9.

$3^x = 9$, hence $x = 2$. The order of the reaction with respect to NO is 2. Now, we can write the rate law.

$$\text{Rate} = k[NO]^2[O_2]$$

We can determine the value of k by choosing any one of the trials and substituting the known values for the concentrations and rate. Inserting the values from trial 1 into the rate law yields

$$1.2 \times 10^{-8}\,M/s = k(0.10)^2(0.10), \text{ and solving for } k \text{ yields}$$
$$k = 1.2 \times 10^{-5}\,M^{-2}s^{-1}.$$

Thus, the rate law for this reaction is:

$$\text{Rate} = (1.2 \times 10^{-5}\ M^{-2}s^{-1})[NO]^2[O_2].$$

As long as the reaction occurs at the temperature for which this rate law was determined, the rate can be determined by plugging in the initial concentrations of the reactants. The value of k changes with temperature, so this k value is only true at the specific temperature for which the data was determined.

In a number of reactions, the order of the reaction for a particular reactant will be determined to be zero. This indicates that the reaction rate does not depend on the concentration of that reactant and the reactant will not appear in the rate law. (Anything raised to the power of 0 equals 1.)

Consider the following reaction and experimental data.

$$A + B + C \rightarrow \text{Products}$$

TABLE 23.2: Experimental Rate Data Table

Trial	Initial [A]	Initial [B]	Initial [C]	Rate
1	1.0 M	1.0 M	1.0 M	0.40 M/s
2	1.0 M	1.0 M	2.0 M	0.40 M/s
3	1.0 M	2.0 M	1.0 M	1.6 M/s
4	2.0 M	2.0 M	1.0 M	1.6 M/s

The reaction rate will be related to the equation $R = k[A]^a[B]^b[C]^c$

Comparing trials 1 and 2, we have [A] and [B] remaining constant while [C] is doubled. The rate also remains the same.

$$(\text{multiple of C concentration})^{\text{c}} = (\text{multiple of rate})$$

$2^c = 1$, so the exponent, c, must equal 0 anything to the power of zero equals 1.

Comparing trials 3 and 4, we have [B] and [C] remaining constant while [A] is doubled. The rate remains the same.

$$(\text{multiple of A concentration})^{\text{a}} = (\text{multiple of rate})$$

$2^a =$ 1, so the exponent, a, must equal 0 anything to the power of zero equals 1.

Comparing trials 1 and 3, we have [A] and [C] remaining constant while [B] is doubled. The rate increases by a factor of 4.

$$(\text{multiple of B concentration})^{\text{b}} = (\text{multiple of rate})$$

$2^b = 4$, so the exponent, b, must equal 2.

Therefore, the rate expression will be:

$$\text{Rate} = k[A]^0[B]^2[C]^0 = k[B]^2.$$

Pacing the Lessons

Use the table below as a guide for the time required to teach the lessons of *Chemical Kinetics*.

TABLE 23.3: 60 Minute Class Periods per Lesson

Lesson	Number of Class Periods
23.1 *Rate of Reactions*	1.5
23.2 *Collision Theory*	1.0
23.3 *Potential Energy Diagrams*	1.0
23.4 *Factors That Affect Reaction Rates*	2.0
23.5 *Reaction Mechanisms*	1.0

Managing Materials

The following items are needed to teach the strategies and activities described in the Teachers Edition of the FlexBook for *Chemical Kinetics*.

TABLE 23.4: Chemical Kinetics Materials List

Lesson	Strategy or Activity	Materials Needed
18.1		
18.2		
18.3		
18.4		
18.5		

Multimedia Resources

Possible Misconceptions

Making the FlexBook Flexible

An important advantage of the FlexBook is the ability it gives you, the teacher, to select the chapters and lessons that you think are most important for your own classes. You should also consult the standards correlation table that follows when selecting chapters and lessons to include in the FlexBook for your classes.

TABLE 23.5: Standard Addressed by the Lessons in Chemical Kinetics

Lesson	California Standards	NSES Standards	AAAS Benchmarks
18.1	8a		
18.2	8b, 8d		
18.3	7b, 8d		
18.4	8b, 8c, 8d		

TABLE 23.5: (continued)

Lesson	California Standards	NSES Standards	AAAS Benchmarks
18.5	8d		

23.1 Rate of Reactions

Key Concepts

In this lesson students explore the rates of chemical reactions.

Lesson Objectives

The student will:

- define chemical kinetics and rates of reactions.
- write the rate expression and the units for the rate expression.
- eefine instantaneous rate.
- calculate instantaneous rate using a tangent line.

Lesson Vocabulary

chemical kinetics The study of rates of chemical reactions and how factors affect rates of reactions.

rate of reaction The measure at which the products are formed over a time interval or the rate at which the reactants are consumed over a time interval.

instantaneous rate The rate of change at a particular time interval.

Strategies to Engage

- Give examples of fast (striking a match), slow (production of coal), and moderate (food spoilage) chemical reactions. Facilitate a discussion where students list factors that affect how fast reactions occur.

Strategies to Explore

Strategies to Extend and Evaluate

- Have each student record the four sentences in this section that most clearly represent the main ideas. Read key sentences in the text and have students raise their hands if they have recorded that sentence. Facilitate a discussion in which students defend their selections.

Lesson Worksheets

There are no worksheets for this lesson.

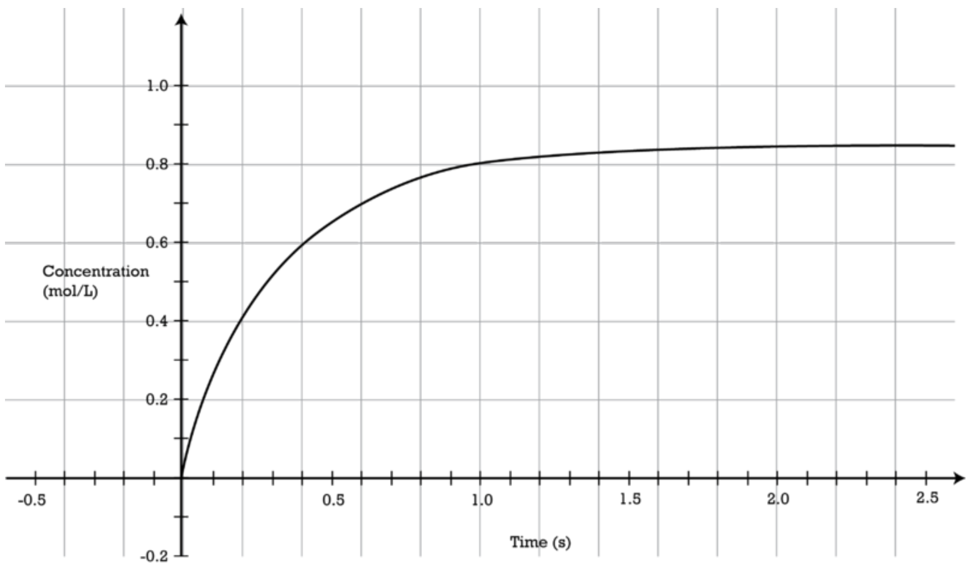
Review Questions

Have students answer the Lesson 18.1 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Given that the concentration of $NO_{2(g)}$ is 0.40 mol/L at 45 s and 0.85 mol/L at 80 s, what is the rate of production of $NO_{2(g)}$ in: $NO_{2(g)} + CO_{(g)} \rightarrow NO_{(g)} + CO_{2(g)}$?
2. For the graph below, draw a tangent line at $t = 0.40$ s and calculate the instantaneous rate.



3. Which expression represents the rate for the product formation for the reaction: $Mg_{(s)} + 2\,HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$?
   a. rate $= \frac{\triangle[Mg]}{\triangle t}$
   b. rate $= \frac{\triangle[HCl]}{\triangle t}$
   c. rate $= \frac{\triangle[MgCl_2]}{\triangle t}$
   d. All of these are accurate representations of the rate.
4. Which statement represents a rate?
   a. The speed of a car is 50 km/h.
   b. Half the product is produced.
   c. A family consumes 5 L of milk.
   d. I ran for 45 minutes.
5. Which statement about the instantaneous rate of a reaction is correct?
   a. The higher the rate, the smaller the slope of a line on a concentration-time graph.
   b. The instantaneous rate is the slope of the tangent to a line on a concentration-time graph.
   c. The instantaneous rate is the slope of the cosine to a line on a concentration-time graph.

d. All of these statements are correct.

6. What is the rate of production of NO gas if the concentration decreases from 0.32 mol/L at 56 s and 0.94 mol/L at 78 s for the reaction $4\ NH_{3(g)} + 5\ O_{2(g)} \rightarrow 4\ NO_{(g)} + 6\ H_2O_{(g)}$?
   a. $-35\ mol/L \cdot s$
   b. $-2.8 \times 10^2\ mol/L \cdot s$
   c. $2.8 \times 10^{-2}\ mol/L \cdot s$
   d. $35\ mol/L \cdot s$
7. It takes 15 minutes for the concentration of a reactant to decrease from 0.45 mol/L to 0.030 mol/L. What is the rate of reaction in $mol/L \cdot s$?

23.2 Collision Theory

Key Concepts

In this lesson students explore reactions rates in terms of collisions between reacting particles.

Lesson Objectives

The student will:

- define the collision theory.
- describe the conditions for successful collisions.
- explain how the kinetic molecular theory applies to the collision theory.
- describe the rate in terms of the conditions of successful collisions.

Lesson Vocabulary

collision theory Explains why reactions occur at this particle level between atoms, ions, and/or molecules. More importantly from the collision theory, is the ability to predict what conditions are necessary for a successful reaction to take place.

kinetic molecular theory Provides the foundation for the collision theory on the atomic level. The collisions between particles are considered to elastic in nature.

threshold energy The minimum amount of energy necessary for a reaction to take place.

collision frequency The total number of collisions per second.

Strategies to Engage

- Begin the lesson by reviewing the kinetic molecular theory with students.

Strategies to Explore

- Have students write a paragraph explaining Figure 18.4 that includes using this lesson's vocabulary terms.

Strategies to Extend and Evaluate

- Have students create a concept map relating the terms and objectives of the concepts explored so far in this chapter.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 18.2 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. According to the collision theory, it is not enough for particles to collide in order to have a successful reaction to produce products. Explain
2. Due to the number of requirements for a successful collision, according to the collision theory, the percentage of successful collisions is extremely small. Yet, chemical reactions are still observed at room temperature and some at very reasonable rates. Explain.
3. What is a basic assumption of the kinetic molecular theory?
   a. All particles will lose energy as the velocity increases
   b. All particles will lose energy as the temperature increases
   c. All particles will increase velocity as the temperature decreases
   d. All particles are in random motion
4. According to the collision theory, which of the following must happen in order for a reaction to be successful: i. particles must collide, ii. particles must have proper geometric orientation, iii. particles must have collisions with enough energy?
   a. i, ii
   b. i, iii
   c. ii, iii
   d. i, ii, iii
5. What would happen in a collision between two particles if there was insufficient kinetic energy and improper geometric orientation?
   a. The particles would rebound and there would be no reaction.
   b. The particles would keep bouncing off each other until they eventually react, therefore the rate would be slow.
   c. The particles would still collide but the byproducts would form.
   d. The temperature of the reaction vessel would increase.
6. Illustrate the successful collision that would occur between the following: $2\ NO + 2\ H_2 \rightarrow N_2 + 2\ H_2O$.

23.3 Potential Energy Diagrams

Key Concepts

In this lesson students learn what information is contained in potential energy diagrams for endothermic and exothermic reactions, how to read them, and how to draw them.

Lesson Objectives

The student will:

- define enthalpy, activation energy, activated complex.
- describe and draw the difference between endothermic and exothermic potential energy diagrams.
- draw and label the parts of a potential energy diagram.

Lesson Vocabulary

potential energy diagrams Potential energy diagrams in the study of kinetics show how the potential energy changes during reactions from reactants to products.

potential energy The potential energy measures the energy stored within the bonds of the reactants and products, and therefore is the internal energy.

exothermic reactions Reactions that have a potential energy difference between the products and reactants that is negative.

endothermic reactions Reactions that have a potential energy difference between the products and reactants that is positive.

activation energy The minimum amount of energy that needs to be supplied to the system so that a reaction can occur.

activated complex A high energy transitional state between the reactants and products.

Strategies to Engage

Strategies to Explore

- Have students write down the lesson objectives, leaving about 5 or 6 lines of space in between. As you explore the lesson, have students write the "answer" to each objective.

Strategies to Extend and Evaluate

- Have students play a review game called "Two Truths and a Lie" using what they know about potential energy diagrams. To do this, pair students, and have each pair write three statements, two of which are facts about potential energy diagrams, and one of which is a plausible "lie." Then have each pair join with two other pairs to share what they wrote, and try to guess which of the statements are "lies" and which are "truths."

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Potential Energy Diagrams**. Ask students to complete the worksheet alone or in pairs as a review of lesson content.

Review Questions

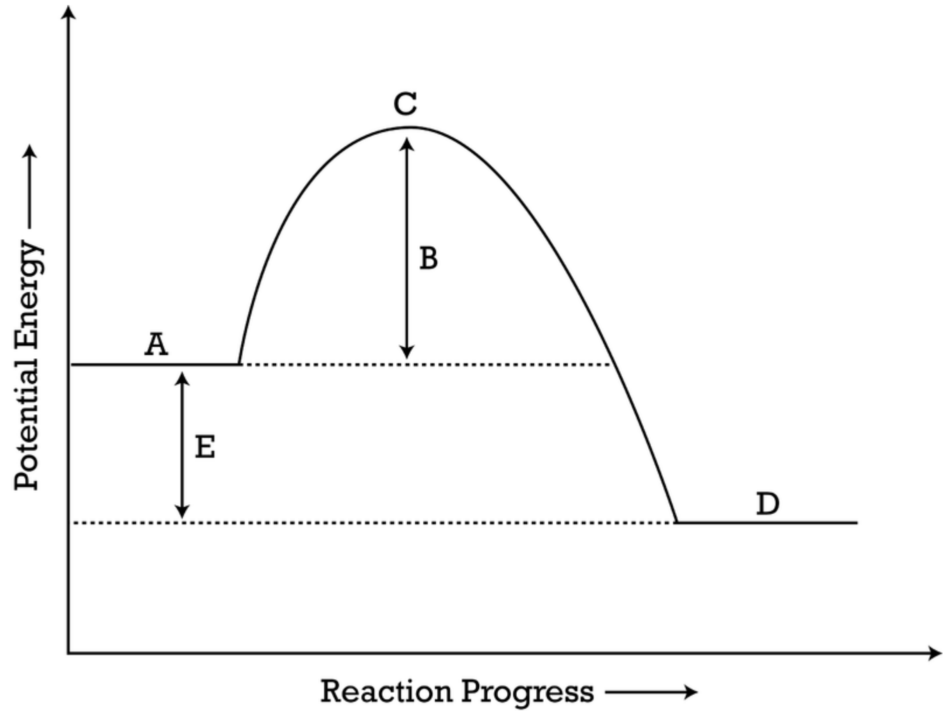
Have students answer the Lesson 18.3 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Define and explain the importance of the activation energy.

Use the diagram below to answer questions 2 through 6.



2. Which letter represents the activation energy barrier?
   a. a
   b. b
   c. c
   d. d
3. Which statement best describes the reaction?
   a. The reaction is exothermic in the forward reaction.
   b. The reaction is endothermic in the forward reaction.
   c. The reaction is exothermic in the reverse reaction.
   d. The reaction is exothermic only at high temperatures.
4. Which letter represents the change in enthalpy for the reaction?
   a. b
   b. c
   c. d
   d. e
5. Which letter represents the activated complex for the reaction?
   a. a
   b. b
   c. c
   d. d
6. What is an activated complex?
   a. a transitional species that can eventually be isolated
   b. a transitional species of that must be made before the products can be formed
   c. a reactant molecule breaking into a product molecule
   d. part of the activation energy barrier
7. For the following reaction, the activation energy is 60 kJ: $A_{2(g)} + 2\ B_{(g)} \rightarrow 2\ AB_{(g)}$. Draw a potential energy diagram properly labeling the following:

a. the axes
b. the reactants and products
c. the activation energy
d. the enthalpy

23.4 Factors That Affect Reaction Rates

Key Concepts

In this lesson students explore the effect of temperature, surface area, concentration, and catalysts on the rate of a chemical reaction.

Lesson Objectives

The student will:

- state how the rate of reaction changes as a function of temperature.
- explain how increased temperature increases the number of particles that can overcome the energy barrier.
- describe the effect of increasing the concentration on the rate of a reaction.
- indicate which reactants in a multi-step process can affect the rate of a reaction.
- calculate, using experimental data, the relationship between the ratio of the change in concentration of reactants, and ratio of the change in rate.
- describe the surface area to volume ratio.
- describe the effect of surface area on reaction rate.
- describe how the change in the surface area affects the collision frequency.
- describe real world examples of the effect of surface area on reaction rate.
- define a catalyst.
- identify a catalyst in a single equation.
- identify a catalyst in a multi-step process.
- describe how a catalyst affects the potential energy diagram.
- explain how a catalyst affects the rate of the reaction.
- explain how a catalyst affects our everyday lives, particularly with vitamins.

Lesson Vocabulary

effective collision A collision that results in a reaction.

multi-step process Reactions that take more than one step in order to make the products.

surface area to volume ratio The comparison of the volume inside a solid to the area exposed on the surface.

catalyst A substance that speeds up the rate of the reaction without itself being consumed by the reaction.

Strategies to Engage

- Perform the ***Catalytic Oxidation of Alcohol with Copper*** *and* **Hydrogen Peroxide Catalyzed by Manganese Dioxide** demonstrations. These demonstrations are located in the Supplemental Lab Book.

Strategies to Explore

- This lesson includes a description of the factors that affect reaction rates. Before reading, prepare less proficient readers by having students write the following on the top of separate sheets of notebook paper:

The Nature of the Reactants

Temperature

Concentration

Surface Area

Catalyst

As they read each section, have them write key points under each heading. This will give the students a quick reference and help them to organize the information. Instruct students to write a one-paragraph summary of the information they have read in each section. DI **Less Proficient Readers**

- Have students complete the labs ***Factors Affecting Reaction Rates*** and **'The Iodine Clock Reaction.** These labs are located in the Supplemental Lab Book.

Strategies to Extend and Evaluate

- Have students research real-world examples of the effect of temperature, surface area, concentration, and catalysts on the rate of chemical reactions. Students should be prepared to discuss their findings with the class.

- Have students read the ***Catalytic Converters*** extra reading located in this book.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Lesson 4 contains four sub-lessons and each sub-lesson has a set of review questions. Have the students answer the Review Questions for each sub-lesson as you cover the sub-lesson.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Why does an increase in temperature increase the rate of the reaction?
2. Why does higher temperature increase the reaction rate?
   a. more of the reacting molecules will have higher kinetic energy
   b. increasing the temperature causes the reactant molecules to heat up
   c. the activation energy will decrease
   d. increasing the temperature causes the potential energy to decrease
3. When the temperature is increased, what does not change?
   a. number of collisions
   b. activation energy requirement
   c. number of successful collisions
   d. all of the above change
4. What is the rule of thumb used for the temperature dependence on the rate?
5. The rule of thumb for the temperature effect on reaction rates is that a reaction rate will double for each 10°C rise in temperature. The rate of reaction for a hypothetical reaction was found to be 0.62 $\text{mol/L} \cdot \text{s}$ at 6°C. What would be the rate at 46°C?
6. Explain how concentration affects reaction rate using the collision theory. You may want to include a diagram to help illustrate your explanation.
7. Why is the increase in concentration directly proportional to the rate of the reaction?
   a. The kinetic energy increases.
   b. The activation energy increases.
   c. The number of successful collisions increases.
   d. All of the above.
8. For the reaction $H_{2(g)} + Cl_{2(g)} \rightarrow 2\ HCl_{(g)}$, an experiment shows that if the concentration of $H_{2(g)}$ is doubled, the rate of reaction stays the same. If the concentration of $Cl_{2(g)}$ doubles, the rate of the reaction quadruples. What is the explanation for this observation?
   a. The reaction is nearing completion and all $H_{2(g)}$ is used up.
   b. The reaction occurs in more than one step.
   c. Excess $Cl_{2(g)}$ has been added.
   d. Not enough information is given.
9. The mechanism for a reaction is as follows:

$$\begin{array}{lll} NO + Br_2 \rightarrow NOBr_2 & \text{(reaction 1)} & \text{(slow)} \\ NOBr_2 + NO \rightarrow 2\ NOBr & \text{(reaction 2)} & \text{(fast)} \end{array}$$

   Which of the following would have the greatest effect on the rate of reaction?
   a. Increase [NO]
   b. Increase $[Br_2]$
   c. Increase $[NOBr_2]$
   d. Increase [NO] and $[Br_2]$
10. Consider the following reaction mechanism. For which substance would a change in concentration have the greatest effect on the rate of the overall reaction?

$$\begin{array}{ll} 2\,A \rightarrow B + 2\,C & \text{(slow)} \\ B + C \rightarrow D + E & \text{(fast)} \end{array}$$

   a. A, B, C

b. A
c. B
d. C

11. The following data (**Table** 23.6) were obtained for the decomposition of N_2O_5 in $CCl_{4(aq)}$ at 45°C. Determine the effect of decreasing the $[N_2O_5]$ on the rate of the reaction.

$$\frac{[N_2O_5]_{\text{trial 2}}}{[N_2O_5]_{\text{trial 1}}} = \frac{0.274}{0.316} = 0.867$$
$$\frac{\text{rate}_{\text{trial 2}}}{\text{rate}_{\text{trial 1}}} = \frac{0.34}{0.39} = 0.87$$

TABLE 23.6:

Trial	N_2O_5 mol/L	Rate$(\text{mol}/\text{L}\cdot s)$
1	0.316	0.39
2	0.274	0.34
3	0.238	0.29
4	0.206	0.25
5	0.179	0.22

12. Why, using the collision theory, do reactions with higher surface area have faster reaction rates?
13. When does an increase in surface area not increase the rate of reaction?
    a. The rate will not be increased if there is insufficient activation energy present.
    b. The rate will not increase if there is not an increase in collisions.
    c. The rate will not increase if the concentration doesn't change.
    d. The rate will not increase if $\triangle H$ does not increase.
14. Choose the substance with the greatest surface in the following groupings:
    a. a block of ice or crushed ice
    b. sugar cubes or sugar crystals
    c. a piece of wood or wood shavings
    d. $O_{2(s)}$ or $O_{2(g)}$
    e. $AgNO_{3(s)}$ or $AgNO_{3(aq)}$
15. Lighter fluid is sometimes used to get a barbecue coals to begin to burn. Give a complete explanation for
    a. the purpose of the lighter fluid; and,
    b. the purpose of the coals.
16. Draw a potential energy diagram for an exothermic reaction labeling the following:
    a. the activation energy of 125 kJ
    b. the enthalpy of -85 kJ/mol
    c. the reactants and product
    d. the axes
    e. the activation energy for the catalyzed reaction
17. The main function of a catalyst is to
    a. provide an alternate reaction pathway
    b. change the kinetic energy of the reacting particles
    c. eliminate the slow step

   d. add another reactant

18. What happens when a catalyst is added?
   a. the activation energy of the forward reaction is lowered
   b. the activation energy of the reverse reaction is lowered
   c. the activation energy in general is lowered
   d. the enthalpy of the reaction is lowered

19. Given the reaction mechanism shown below, which species is the catalyst?

$COCl_{2(g)} \rightarrow COCl_{(g)} + Cl_{(g)}$	(fast)	(reaction 1)
$Cl_{(g)} + COCl_{2(g)} \rightarrow COCl_{(g)} + Cl_{2(g)}$	(slow)	(reaction 2)
$2\ COCl_{(g)} \rightarrow 2\ CO_{(g)} + 2\ Cl_{(g)}$	(fast)	(reaction 3)
$2\ Cl_{(g)} \rightarrow Cl_{2(g)}$	(fast)	(reaction 4)

   a. $CO_{(g)}$
   b. $COCl_{2(g)}$
   c. $COCl_{(g)}$
   d. $Cl_{(g)}$

20. Catalysts are used in all parts of society from inside our bodies to the largest industries in the world. Give an example of a catalyst and explain its usefulness.

23.5 Reaction Mechanism

Key Concepts

In this lesson students explore multi-step processes as well as the individual reactions in the multi-step process.

Lesson Objectives

The student will:

- define reaction mechanisms.
- identify the rate-determining step.
- draw a potential energy diagram for a multi-step process.

Lesson Vocabulary

elementary step A single, simple step in a multi-step process involving one or two particles.

reaction mechanism Most reactions do not take place in one step but rather occur as a combination of two or more elementary steps.

rate-determining step The slowest step in the reaction mechanism.

Strategies to Engage

- Preview the lesson vocabulary to find out what your students already know about the concepts to be explored in this lesson. Have students define each vocabulary term. At the end of the lesson encourage students to go back and write the correct definition for each incorrect definition.

Strategies to Explore

- Have students write a paragraph explaining Figure 18.12 that includes using this lesson's vocabulary terms.

Strategies to Extend and Evaluate

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 18.5 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Why do most reactions take place in more than one step?
2. The overall rate of a reaction depends on
   a. the temperature
   b. the surface area
   c. the pressure
   d. the slowest step
3. Suppose a reaction takes place according to the following reaction mechanism. Determine which step in the mechanism is the rate-determining step.
4. If you wanted to increase the overall rate of the reaction in Question #2, would increase the concentration of X or Y? Explain

$X+Y \rightarrow Z$	(fast)
$X+Z \rightarrow F$	(slow)

4. Consider the following equation for the formation of ammonia. Explain why this equation is not likely to represent the reaction mechanism.

Chapter 23 Enrichment

Extra Readings

Catalytic Converters

In 2007, there were approximately 200 *million* vehicles on American roads, with perhaps 450 *million* operating in the world. For each of these cars and trucks, their mobility stems from an internal combustion engine. In these engines, hydrocarbons combine with oxygen in a combustion process, producing carbon dioxide and water vapor gases, in addition to a host of other combustion byproducts, such as nitrogen oxides and sulfur oxides. Due to

their sheer numbers and the volume of gases produced, environmental concerns about increasing levels of carbon dioxide emissions leading to global warming, and the links of nitrogen and sulfur oxides with acid rain, have led to the development of catalytic convertors, first implemented in automobiles in 1975. These catalytic convertors are devices installed in the emission train of automobiles containing catalysts consisting of precious metal surfaces that convert harmful emission gases like NO_x and CO into N_2 and CO_2 respectively.

Catalysts are substances that can enable or accelerate chemical reactions without being consumed in the reactions themselves. In the case of catalytic convertors, originally developed by Eugene Houdry to minimize the effects of automobile exhaust, the system consists of a large surface area porous support for the actual catalyst, a precious or rare metal oxide, like those of platinum, palladium or rhodium. The amount of metals due to their extreme cost is minimal, usually consisting of a thin surface on the support.

Modern catalytic convertors are usually formulated to be three way systems: 1. reducing NO_x to N_2; 2. oxidizing carbon monoxide and 3. oxidizing any remaining incombustible hydrocarbons to carbon dioxide. Emission gases exiting from the engine reach the exhaust manifold at temperatures about $500°C$. As the chemical reductions and oxidations occur only on the surface, only a small amount, about $\frac{1}{5}$ of an ounce of palladium metal is needed. Its enormous effectiveness in this capacity outstrips the extreme cost, currently more than $200 per ounce leads to two additional problems: how to recover the metal when the automobile is no longer in use, and more pressingly, the issue of catalytic converter theft to recover the precious metal. Car enthusiasts also lament that the use of the catalytic converter in the exhaust system compromises the horsepower output and back pressure yet the Federal Clean Air Act mandates their inclusion in all vehicles.

The catalytic converter is a modern–day environmental success story, removing hundreds of tons of carbon monoxide, and 30[U+0080][U+0093]50 *million* tons of excess hydrocarbons and nitrogen oxides. Air quality in the United States, even with a growing population and ever–increasing dependence on fossil fuel vehicles, is vastly improved in all major categories since 1975.

One problem that may exist with catalytic converters is that in recent years, researchers at the University of California have detected significantly higher levels of another pollutant, ammonia, in the exhaust of automobiles with catalytic converters. The researchers indicate that while catalytic converters have played a major part in reducing air pollution caused by automobiles, these latest findings suggests that while fixing one problem, converters may have caused an unexpected secondary problem. Further research will be necessary to determine if the catalytic converters are producing the higher levels of ammonia in automobile exhaust.

CHAPTER 24 Chemical Equilibrium TE

Chapter Outline

Chemical Equilibrium

Outline

This chapter *Chemical Equilibrium* consists of four lessons that cover relationships between forward and reverse reaction rates, the concept of chemical equilibrium, the mathematics of the equilibrium constant, Le Chatelier's principle, and solubility product constant calculations.

- **Lesson 19.1 Introduction to Equilibrium**
- **Lesson 19.2 Equilibrium Constant**
- **Lesson 19.3 The Effect of Applying Stress to Reactions at Equilibrium**
- **Lesson 19.4 Slightly Soluble Salts**

Overview

In these lessons, students will explore:

- The conditions of chemical equilibrium.
- Equilibrium constant expressions.
- Le Chatelier's Principle.
- Solubility product constant expressions.

Science Background Information

This information is provided for teachers who are just beginning to instruct in this subject area.

Chemical Equilibrium

In principle, any reaction that can be represented by a balanced chemical equation can take place. There are, however, two situations which may inhibit the reaction from occurring.

- The thermodynamic tendency (the combination of entropy and enthalpy) for the reaction to occur may be so small that the quantity of products is very low, or even negligible. This type of chemical reaction is said to be ***thermodynamically inhibited***.
- The rate at which the reaction proceeds may be so slow that many years are required to detect any product at all, in which case we say the reaction is ***kinetically inhibited***.

As a reaction proceeds, the quantities of the components on one side of the reaction equation will decrease and those on the other side will increase. As the concentrations of the components on one side of the equation decrease, that reaction rate slows down. As the concentrations of the components on the other side of the equation increase, that reaction rate speeds up. Eventually the two reaction rates become equal and the composition of the system stops changing. At this point, the reaction is in it's *equilibrium state* and no further change in composition will occur, as long as the system is left undisturbed.

In many reactions, the equilibrium state occurs when significant amounts of both reactants and products are present. Such a reaction is said to be *reversible*. The equilibrium composition is independent of the direction from which it is approached. The labeling of substances as *reactants* and *products* is entirely a matter of convenience.

The ***law of mass action*** states that any chemical change is a competition between a forward reaction (left–to-right) and a reverse reaction (right-to-left). The rates of these two reactions are governed by the concentrations of the substances reacting, and the temperature. As the reaction proceeds, these two reaction rates approach each other in magnitude and at equilibrium, they become equal.

Since the reactions continue at equilibrium (at equal rates), equilibrium is referred to as *dynamic* equilibrium. At equilibrium, *microscopic changes* (the forward and reverse reactions) continue but *macroscopic changes* (changes in quantities of substances) cease.

When a chemical system is at equilibrium, any disturbance of the system, such as a change in temperature, or the addition or removal of a reactant or product, will cause the equilibrium to shift to a new equilibrium state (different quantities of reaction components). The disturbance in the system causes changes in the reaction rates and quantities of components change until the reaction rates again become identical, and a new equilibrium position is established.

Pacing the Lessons

Use the table below as a guide for the time required to teach the lessons of *Chemical Equilibrium*.

TABLE 24.1: 60 Minute Class Periods per Lesson

Lesson	Number of Class Periods
19.1 *Introduction to Equilibrium*	1.0
19.2 *Equilibrium Constant*	1.5
19.3 *The Effect of Applying Stress to Reactions at Equilibrium*	2.0
19.4 *Slightly Soluble Salts*	1.5

Managing Materials

The following items are needed to teach the strategies and activities described in the Teachers Edition of the FlexBook for *Chemical Equilibrium*.

TABLE 24.2: **Chemical Equilibrium Materials List**

Lesson	Strategy or Activity	Materials Needed
19.1		
19.2		
19.3		
19.4		
19.5		

Multimedia Resources

You may find these additional web based resources helpful when teaching *Chemical Equilibrium*:

- Le Chatelier's Principle Movie: http://genchemist.wordpress.com/2007/10/05/le-chateliers-principle-the-movie/
- Le Chatelier's Principle and the Haber Process video clip: http://videos.howstuffworks.com/hsw/12468-chemistry-connections-le-chatelier-and-haber-bosch-video.htm

Possible Misconceptions

Identify:

Many students assume that all chemical changes are irreversible.

Clarify:

A reaction that proceeds in only one direction is known as an irreversible reaction. A chemical reaction in which the product(s) can react to produce the original reactant(s) is a reversible reaction. For example, the reaction of calcium oxide, CaO, with carbon dioxide, CO_2, produces calcium carbonate, $CaCO_3$. If the calcium carbonate is heated, the reaction produces calcium oxide and carbon dioxide - the original reactants.

Promote Understanding:

Place about 10 mL of concentrated nitric acid in a beaker. Add a penny. Pour the red gas (NO_2) that results into a test tube and stopper it. Place the test tube into an ice bath. The gas, N_2O_4, will become almost colorless. Return the test tube to room temperature. The gas, NO_2, will return to its red color.

Discuss:

Write the equation: $2\ NO_{2(g)} \leftrightharpoons N_2O_{4(g)}$ on the board. Inform students that the substance on the left side of the arrow is the reactant and the substance on the right side of the arrow is the product of the chemical reaction.

Ask:

"How do you know that a chemical reaction took place?" (A new substance was formed.)

Ask:

How do you know that the reaction is reversible? (The product was able to re-form the original reactant.)

Making the FlexBook Flexible

An important advantage of the FlexBook is the ability it gives you, the teacher, to select the chapters and lessons that you think are most important for your own classes. You should also consult the standards correlation table that follows when selecting chapters and lessons to include in the FlexBook for your classes.

TABLE 24.3: **Standard Addressed by the Lessons in Chemical Equilibrium**

Lesson	California Standards	NSES Standards	AAAS Benchmarks
19.1	9b		
19.2	9c		
19.3	9a		
19.4			

24.1 Introduction to Equilibrium

Key Concepts

In this lesson students explore the conditions of chemical equilibrium.

Lesson Objectives

The student will:

- describe the three possibilities that exist when reactants come together.
- identify the three possibilities by looking at a chemical equation.
- describe what is occurring in a system at equilibrium.
- define dynamic equilibrium.
- define the conditions of dynamic equilibrium.

Lesson Vocabulary

chemical equilibrium A state that occurs when the rate of forward reaction is equal to the rate of the reverse reaction.

dynamic equilibrium A state that continues in which the rate of the forward reaction is equal to the rate of the reverse reaction; or, the number of particles/molecules of the reactant becoming the product is equal to the number of particle/molecules of the product becoming the reactant.

Strategies to Engage

- Explain to students that chemical reactions do not always go completely to products, and that in this lesson, they will explore the three possibilities that exist when reactants come together.
- Explain to students that in this lesson, they will explore the concept of dynamic equilibrium. Give students the opportunity to define this term just by examining the words themselves. At the end of the lesson, have students check their original definition and make corrections, if necessary.

Strategies to Explore

- Have students write down the lesson objectives, leaving about 5 or 6 lines of space in between. As you explore the lesson, have students write the "answer" to each objective.

- Challenge groups of students to create and perform a short skit to demonstrate chemical equilibrium.
- Perform the ***A Light Activated Reversible Reaction*** *demonstration. This demonstration is located in the Supplemental Lab Book.*

Strategies to Extend and Evaluate

- As a review of lesson concepts, make a copy of Figure 19.1. Below it, rewrite the paragraph explaining Figure 1 deleting key words and create an overhead. Have students choose words to fill in these blanks so that the text makes sense.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

There are two sub-lessons in lesson 19.1 that each have a set of review questions. Have the students answer the review questions for each sub-lesson as you cover it.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. What does the term dynamic equilibrium mean?
2. List all of the conditions for a dynamic equilibrium.
3. Of the following conditions, which is not required for a dynamic equilibrium?
    a. rate of the forward reaction equals the rate of the reverse reaction.
    b. reaction occurs in an open system
    c. reaction occurs at a constant temperature
    d. reaction occurs in a closed system
4. Which of the following systems, at room temperature and pressure, can be described as a dynamic equilibrium?
    a. an open flask containing air, water and water vapor
    b. a glass of water containing ice cube cubes and cold water
    c. a closed bottle of soda pop
    d. an open flask containing solid naphthalene
5. Is each of the following in a state of equilibrium? Explain.
    a. Ice cubes are melting in a glass of water with a lid on it
    b. Crystals of potassium dichromate were dissolved in water, and now the water is a uniform orange color with a small amount of crystal left in the closed container.
    c. An apple that is left on the counter for a few days, it dries out and turns brown.
6. If the following table (**Table** 24.4) of concentration vs. time was provided to you for the ionization of acetic acid, how would you know when equilibrium was reached?

TABLE 24.4: Data Table For Problem 6

Time (min)	$[HC_2H_3O_2]$ mol/L
0	0.100
0.5	0.099
1.0	0.098
1.5	0.097
2.0	0.096
2.5	0.095
3.0	0.095
3.5	0.095
4.0	0.095
4.5	0.095
5.0	0.095

24.2 Equilibrium Constant

Key Concepts

In this lesson students explore equilibrium constant expressions.

Lesson Objectives

The student will:

- write equilibrium constant expressions.
- use equilibrium constant expressions to solve for unknown concentrations.
- use known concentrations to solve for the equilibrium constants.
- explain what the value of K means in terms of relative concentrations of reactants and products.

Lesson Vocabulary

equilibrium constant A mathematical ratio that shows the concentrations of the products divided by concentration of the reactants.

Strategies to Engage

- Have students read the review questions at the end of this section. This way, students will be familiar with the types of information that they will explore in this section.

Strategies to Explore

- Use the mathematical calculations in this lesson to reduce the reliance on language skills. As you go through each example problem, use them to explain the concepts explored in this lesson. DI **English Language Learners**

Strategies to Extend and Evaluate

- As a review of lesson content, have students write questions derived from Bloom's Taxonomy. Instruct students to research Bloom's Taxonomy and write and answer one question from each of the six levels:

knowledge, comprehension, application, analysis, synthesis, and evaluation.

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Equilibrium**. Ask students to complete the worksheet alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 19.2 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Why are solids and liquids not included in the equilibrium constant expression?
2. What does the value of K mean in terms of the amount of reactants and products?
3. What is the correct equilibrium constant expression for the following reaction: $2\ SO_{2(g)} + O_{2(g)} \rightleftharpoons 2\ SO_{3(g)}$?
   a. $K = \frac{[SO_2]^2[O_2]}{[SO_3]^2}$
   b. $K = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$
   c. $K = \frac{[2\ SO_2][O_2]}{[2\ SO_3]}$
   d. $K = \frac{[2\ SO_3]}{[2\ SO_2][O_2]}$
4. What is the correct equilibrium constant expression for the following reaction: $Cu(OH)_{2(s)} \rightleftharpoons Cu^{2+}_{(aq)} + 2\ OH^-_{(aq)}$?
   a. $K = \frac{[Cu_{2+}][OH_-]^2}{[Cu(OH)^2]}$
   b. $K = \frac{[Cu(OH)^2]}{[Cu^{2+}][OH^-]^2}$
   c. $K = \frac{1}{[Cu^{2+}][OH^-]^2}$
   d. $K = [Cu^2+][OH^-]^2$
5. Consider the following equilibrium system: $2\ NO_{(g)} + Cl_{2(g)} \rightleftharpoons 2\ NOCl_{(g)}$. At a certain temperature, the equilibrium concentrations are as follows: $[NO] = 0.184\ mol/L$, $[Cl_2] = 0.165\ mol/L$, and $[NOCl] = 0.060\ mol/L$. What is the equilibrium constant for this reaction?
   a. 0.506
   b. 0.648
   c. 1.55
   d. 1.97
6. For the reaction $2\ MgCl_{2(s)} + O_{2(g)} \rightleftharpoons 2\ MgO_{(s)} + 2\ Cl_{2(g)}$, the equilibrium constant was found to be 3.86 at a certain temperature. If 0.560 mol $O_{2(g)}$ is placed in a 1.00 L container, what is the concentration of $Cl_{2(g)}$ at equilibrium?
   a. 1.47 mol/L
   b. 2.16 mol/L
   c. 2.88 mol/L
   d. not enough information is available
7. Write the equilibrium constant expressions for each of the following equations:

a. $CH_3NH_{2(aq)} + H_2O_{(l)} \rightleftharpoons CH_3NH^+_{3(aq)} + OH^-_{(aq)}$
b. $2\ CaSO_{4(s)} \rightleftharpoons 2\ CaO_{(s)} + 2\ SO_{2(g)} + O_{2(g)}$
c. $2\ Fe^{3+}_{(aq)} + 3\ S^{2-}_{(aq)} \rightleftharpoons Fe_2S_{3(s)}$
d. $Hg_{(l)} + H_2S_{(g)} \rightleftharpoons HgS_{(s)} + H_{2(g)}$

24.3 The Effect of Applying Stress to Reactions at Equilibrium

Key Concepts

In this lesson students explore Le Châtelier's Principle.

Lesson Objectives

The student will:

- state Le Châtelier's Principle.
- demonstrate on specified chemical reactions how Le Châtelier's Principle is applied to equilibrium systems.
- describe the effect of concentration on an equilibrium system.
- demonstrate with specific equations how Le Châtelier's Principle explains the effect of concentration.
- describe the effect of pressure as a stress on the equilibrium position.
- describe the pressure effect in Le Châtelier's Principle.
- describe the effect of temperature as a stress on an equilibrium system.
- explain how Le Châtelier's principle explains the effect of temperature.
- explain how a catalyst works in equilibrium reactions.
- explain the effect of a catalyst in equilibrium positions.

Lesson Vocabulary

Le Châtelier's Principle Applying a stress to an equilibrium system causes the equilibrium position to shift to offset that stress and regain equilibrium.

The Haber Process A commercial method that makes the maximum amount of ammonia using the Le Châtelier's Principle.

exothermic reaction A reaction in which the heat content of the reactants is greater than the heat content of the products. The excess energy is given off as a product.

endothermic reaction A reaction in which the heat content of the reactants is less than the heat content of the products. Energy is needed to be added to the reactants in order to form the products.

catalyst A substance that increases the rate of a chemical reaction but is, itself, left unchanged, at the end of the reaction.

Strategies to Engage

- Before beginning this lesson, review with students reversible reactions and the concept of chemical equilibrium.
- Perform the ***Equilibrium between*** NO_2***and*** N_2O_4*demonstration. This demonstration is located in the Supplemental Lab Book.*

Strategies to Explore

This lesson includes a description of the effects of applying stress to reactions at equilibrium. Before reading, prepare less proficient readers by having students write the following on the top of separate sheets of notebook paper:

- Effect of Concentration Changes
- Effect of Pressure Changes
- Effect of Temperature Changes
- Effect of a Catalyst

As they read each section have them write key points under each heading. This will give the students a quick reference and help them to organize the information. Instruct students to write a one-paragraph summary of the information they have read in each section. DI **Less Proficient Readers**

Strategies to Extend and Evaluate

- Have interested students research practical application of the Le Châtelier's Principle such as the Contact Process, cola drinks, and carbon monoxide poisoning. Students should be prepared to share their findings with their classmates.

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Le Châtelier's Principle**. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Review Questions

Lesson 19.3 has four sub-lessons and each sub-lesson has its own set of Review Questions. Have the students answer the Review Questions for each sub-lesson as you cover it.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

24.4 Slightly Soluble Salts

Key Concepts

In this lesson students explore solubility product constant expressions.

Lesson Objectives

The student will:

- define solubility product constants.
- write solubility product constant expressions.
- calculate solubility product constants.

Lesson Vocabulary

solubility product constant, Equilibrium constant for a slightly soluble salt.

Strategies to Engage

- Before beginning this lesson, review with students equilibrium constants. Explain to students that in this lesson, they will explore equilibrium constants for slightly soluble salts.

Strategies to Explore

- Use the mathematical calculations in this lesson to reduce the reliance on language skills. As you go through each example problem, use them to explain the concepts explored in this lesson. DI **English Language Learners**

Strategies to Extend and Evaluate

Lesson Worksheets

Copy and distribute the lesson worksheets in the *CK-12 Chemistry Workbook* titled **Solubility and K_{sp}. Ask students to complete the worksheets alone or in pairs as a review of lesson content.**

Review Questions

Have students answer the Lesson 19.4 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. What is the solubility product constant? Give an example.
2. Why is solubility considered a special case for chemical equilibria?
3. Nickel hydroxide is a slightly soluble salt. Its dissociation reaction is represented as: $Ni(OH)_{2(s)} \rightleftharpoons Ni^{2+}_{(aq)} + 2\ OH^-_{(aq)}$. Which of the following represents the solubility product constant expression, K_{sp}?
   a. $K_{sp} = \frac{[Ni^{2+}][OH^-]}{[Ni(OH)_2]}$
   b. $K_{sp} = \frac{[Ni^{2+}][OH^-]^2}{[Ni(OH)_2]}$
   c. $K_{sp} = [Ni^{2+}][OH^-]$
   d. $K_{sp} = [Ni^{2+}][OH^-]^2$
4. The K_{sp} for AgBr is 5.0×10^{-13}. What is $[Ag^+]$ at equilibrium?
   a. 5.0×10^{-13} mol/L
   b. 7.1×10^{-7} mol/L
   c. 2.5×10^{-13} mol/L
   d. not enough information is given
5. The K_{sp} for PbF_2 is 3.60×10^{-8}. What is $[F^-]$ at equilibrium?
   a. 3.60×10^{-8} mol/L
   b. 3.33×10^{-3} mol/L
   c. 4.16×10^{-3} mol/L
   d. 2.08×10^{-3} mol/L
6. Magnesium hydroxide is the acive component in milk of magnesia, a suspension used to cure indigestion. It has an equilibrium constant of 6.3×10^{-10}. Write the dissociation equation and comment on the value of the equilibrium constant.
7. Write the dissociation reactions for the following salts as well as the K_{sp} expressions.
   a. calcium fluoride
   b. chromium(II) carbonate
   c. arsenic(III) sulfide

CHAPTER 25 Acids and Bases TE

Chapter Outline

Unit 9 Chemistry of Acids and Bases

Outline

This unit, Chemistry of Acids and Bases, includes two chapters that explore properties and reactions of acids, bases, salts, water, and buffers.

- **Chapter 25 Acids and Bases**
- **Chapter 26 Water, pH, and Titration**

Overview

Acids and Bases

This chapter includes the definitions of acids and bases, the causes of strong and weak acids and bases, the hydrolysis of salts, and an introduction to pH.

Water, pH***, and Titration***

This chapter covers the mathematics of the dissociation of water, acid-base indicators, acid-base titration, and buffers.

Acids and Bases

Outline

This chapter *Acids and Bases* consists of eight lessons that includes the definitions of acids and bases, the causes of strong and weak acids and bases, the hydrolysis of salts, and an introduction to pH.

- **Lesson 25.1 Arrhenius Acids**
- **Lesson 25.2 Strong and Weak Acids**
- **Lesson 25.3 Arrhenius Bases**
- **Lesson 25.4 Salts**
- **Lesson 25.5** pH
- **Lesson 25.6 Weak Acid/Base Equilibria**
- **Lesson 25.7 Bronsted-Lowry Acids and Bases**
- **Lesson 25.8 Lewis Acids and Bases**

Overview

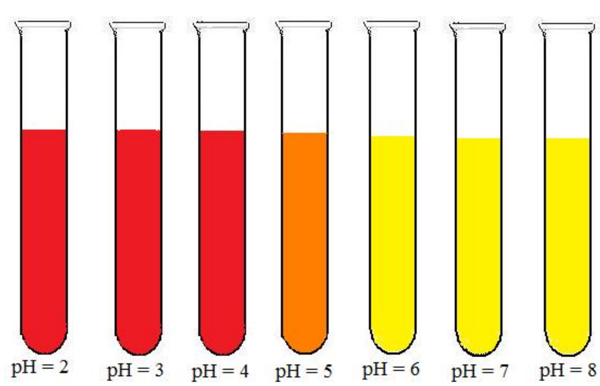
In these lessons, students will explore:

- The Arrhenius acid definition and properties of acids.
- Strong and weak acids in terms of ionization percent.
- The Arrhenius base definition and properties of bases.
- Acid-base neutralization reactions.
- $[H^+]$ and $[OH^-]$ and pH.
- Weak acids and weak bases as equilibrium systems.
- The Bronsted-Lowry definitions of acids and bases, and acid-base conjugate pairs.
- The Lewis definitions of acids and bases, and reactions of Lewis acids and bases.

Science Background Information

The** pH**at which an Indicator Changes Color

Many acid-base indicators exhibit exactly three colors. There is the color of the indicator when it is predominantly in its molecular form, the color of the indicator when it is predominantly in its ionic form, and there is the color of the indicator when it is close to 50% in each form. Consider a fictitious indicator, HIn, whose K_a is 1.0×10^{-5}. At pH values below 5, this indicator is distinctly red, at pH values above 5, it is distinctly yellow, and exactly at a pH of 5, the indicator is orange.



The dissociation equation for the indicator, HIn, is

$$HIn_{(aq)} \leftrightharpoons H^+ + In^-$$

Red Yellow

When the hydrogen ion concentration is high, the equilibrium is shifted toward the reactants, most of the indicator particles are in the form of undissociated molecules, and the solution is red. When the hydrogen ion concentration is low, the equilibrium is shifted toward the products, most of the indicator particles are in the form of anions, and the solution is yellow. At some exact pH, the equilibrium will be adjusted so that exactly 50% of the indicator particles are in the form of undissociated molecules, and 50% in the form of anions. In this case, the solution will be a mixture of equal numbers of red molecules and yellow ions, hence will be orange.

Here is the equilibrium constant expression for the indicator.

$$K_a = \frac{[H^+][In^-]}{[HIn]} = 1.0 \times 10^{-5}$$

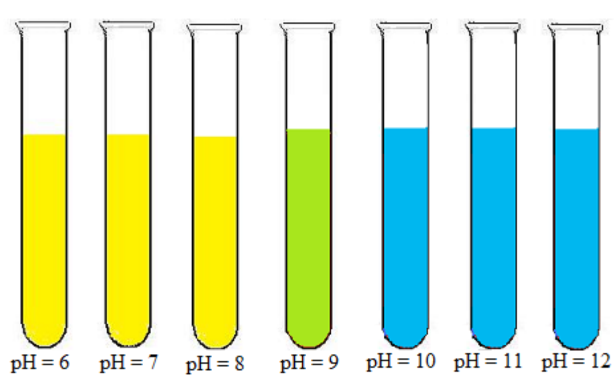
For the pH at which the color changes, we are seeking the point where half of the indicator particles are in each form; in other words, $[In^-] = [HIn]$. When these two values are exactly equal, they will cancel from the expression.

$$K_a = \frac{[H^+][\cancel{In^-}]}{[\cancel{HIn}]} = 1.0 \times 10^{-5}$$

As you can see, mathematically, the $[H^+]$ for this exact point will be equal to the K_a value and $pH = -\log(1.0 \times 10^{-5}) = 5$, which is in agreement with the pictures of the indicator colors at various pH[U+0080][U+0099]s.

Consider the indicator thymol blue. The undissociated molecules of thymol blue are yellow and the anions are blue. The K_a for thymol blue is 1.0×10^{-9}. When this indicator is 50% in the form of undissociated molecules and 50% anions, the $50-50$ mixture of yellow and blue would result in a green color. Calculations of the same type as shown for the previous example indicate that the green color will be present when the $[H^+]$ is equal to the value of the K_a, 1.0×10^{-9}. Therefore, the color change pH for thymol blue is $pH = 9$. When the pH value is less than 9, the indicator will be yellow, at exactly 9, it will be green, and above 9, it will be blue.

It should be clear that putting a few drops of thymol blue in a solution and getting a resultant yellow color does not tell you the pH of the solution. It only tells you that the pH is less than 9. Similarly, a resulting blue solution of thymol blue only tells you that the pH is greater than 9. There is only one color of a thymol blue solution that tells you the pH and that is green.



Pacing the Lessons

Use the table below as a guide for the time required to teach the lessons of *Acids and Bases*.

TABLE 25.1: 60 Minute Class Periods per Lesson

Lesson	Number of Class Periods
25.1 *Arrhenius Acids*	1.0
25.2 *Strong and Weak Acids*	1.0
25.3 *Arrhenius Bases*	1.0
25.4 *Salts*	1.0
25.5 *pH*	1.5
25.6 *Weak Acid/Base Equilibria*	2.0
25.7 *Bronsted-Lowry Acids and Bases*	1.0
25.8 *Lewis Acids and Bases*	1.0

Managing Materials

The following items are needed to teach the strategies and activities described in the Teachers Edition of the FlexBook for *Acids and Bases*.

TABLE 25.2: Acids and Bases Materials List

Lesson	Strategy or Activity	Materials Needed
25.1		
25.2		
25.3		
25.4		
25.5		
25.6		
25.7		
25.8		

Multimedia Resources

You may find these additional web based resources helpful when teaching *Acids and Bases*:

- Interactive pH lab: http://www.proteacher.com/cgi-bin/outsidesite.cgi?id=5268#38;external=http://www.miamisci.org/ph/guide.html#38;original=http://www.proteacher.com/110052.shtml#38;title=The%20pH%20Factor
- pH scale activity: http://www.quia.com/rd/1975.html?AP_rand=201033057
- Titration demonstration:http://www.chem.iastate.edu/group/Greenbowe/sections/projectfolder/flashfiles/stoichiometry/acid_base.html

Possible Misconceptions

Identify: Students may think that acid-base strength is the same as concentration.

Clarify: Concentration refers to the number of moles of solute per liter of solution while strength refers to the degree to which the substance forms ions in solution.

Promote Understanding: Use HCl and $HC_2H_3O_2$ as examples. Explain to students that hydrochloric acid is a strong acid because in water, nearly all of the HCl molecules ionize to form H^+ and Cl^- ions. On the other hand acetic acid is a weak acid because only a small amount of $HC_2H_3O_2$ molecules ionize to form H^+ and $C_2H_3O_2$

[U+0080][U+0093] *ions.Explaintostudentsthatthestrengthofanacidorbasedependsonitsabilitytoionize.Also,ifasolutionhasa*
40 mL *of* 1.0 M *hydrochloricacid,aceticacid,andoxalicacid.Testeachsolutionwithaconductivitytester.*

Discuss: Point out to students that each acid had the same concentration $(1.0\ M)$ but they did not have the same strength. Discuss with students the difference between acid-base strength and concentration.

Making the FlexBook Flexible

An important advantage of the FlexBook is the ability it gives you, the teacher, to select the chapters and lessons that you think are most important for your own classes. You should also consult the standards correlation table that follows when selecting chapters and lessons to include in the FlexBook for your classes.

TABLE 25.3: Standard Addressed by the Lessons in Acids and Bases

Lesson	California Standards	NSES Standards	AAAS Benchmarks
25.1	5a, 5b, 5c		
25.2	5c		
25.3	5a, 5b, 5e		
25.4	5a		
25.5	1e, 5d, 5f		
25.6	1e, 5c		
25.7	5e		
25.8	5e		

25.1 Arrhenius Acids

Key Concepts

In this lesson students explore the Arrhenius acid definition and properties of acids.

Lesson Objectives

- Define an Arrhenius acid and know some examples of acids.
- Define operational and conceptual definition.
- Explain the difference between operational and conceptual definitions.
- Describe the properties of acids.
- Describe some of the reactions that acids undergo.

Lesson Vocabulary

Arrhenius acid A substance that produces H^+ ions in solution.

operational definitions Definitions that describe how something behaves. (i.e. the operational definition of acids includes tastes sour and turns blue litmus red.)

conceptual definitions Definitions that describe why something behaves the way it does. (i.e. the conceptual definition of acids includes reacting with bases to neutralize them.)

Strategies to Engage

- Set up a KWL chart on the board or chart paper. Activate prior knowledge by asking students what they **Know** about acids and bases and write that information in the first column. Have students set goals specifying what they **Want** to learn about acids and bases and write this information in the second column. At the end of the chapter, have students discuss what they **Learned** about acids and bases and write this information in the third column.
- Facilitate a discussion with students about how acids and bases affect our everyday lives. Have students list substances they think contain either acids or bases. Ask students what properties helped them to identify the substances as either acids or bases. Have students come up with their own operational definitions of acids and bases. Use this opportunity to gauge student understanding of acids and bases and to clear up any misconceptions.

Strategies to Explore

Strategies to Extend and Evaluate

- As a review of lesson concepts, have each student record the four sentences in this section that most clearly represent the main ideas. Read key sentences in the text and have students raise their hands if they have recorded that sentence. Facilitate a discussion in which students defend their selections.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 25.1 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Explain the difference between a conceptual definition and an operational definition.
2. What are the properties of acids? Give a common example.
3. Which statement best describes a characteristic of acid solutions?
   a. They react with some metals to form hydrogen gas.
   b. They turn red litmus paper blue.
   c. They taste bitter.
   d. They are made from non-metal oxides.
4. Which of the following is the Arrhenius definition of an acid?
   a. An acid is a substance that donates protons.
   b. An acid is a substance that accepts protons.
   c. An acid is a substance that dissolves in water to form OH^- ions.
   d. An acid is a substance that reacts with water to form H^+ ions.
5. Which of the following will react with acids and produce hydrogen gas?
   a. chlorine
   b. ammonia
   c. carbon
   d. magnesium
6. Write the reaction for each of the following:
   a. hydrofluoric acid + sodium hydroxide
   b. potassium hydroxide + hydrogen sulfide
   c. dissociation of iodic acid
   d. zinc + hydrochloric acid

25.2 Strong and Weak Acids

Key Concepts

In this lesson students explore strong and weak acids in terms of ionization percent.

Lesson Objectives

- Distinguish between strong and weak acids.
- Identify strong and weak acids from given choices.
- Describe how strong and weak acids differ in terms of concentrations of electrolytes.

Lesson Vocabulary

strong acid Acids that completely ionize or undergo 100% ionization in solution (i.e. HCl).

weak acids Acids that do not completely ionize or undergo 100% ionization in solution (i.e. $HC_2H_3O_2$).

Strategies to Engage

- Have students read and propose answers to the questions posed in the lesson introduction. At the end of the lesson have students check their answers, and make the necessary corrections.

Strategies to Explore

- Have students write a paragraph describing how strong and weak acids differ in terms of concentrations of electrolytes.

Strategies to Extend and Evaluate

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Strong Acids and Bases**. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 25.2 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. What is the difference between a strong and weak acid? Show an example of each.
2. In terms of electrolyte solutions, how would you distinguish between a strong acid and a weak acid?
3. All of the following are weak acids except?
   a. $HClO_3$
   b. $HC_2H_3O_2$
   c. HF
   d. HCl
4. Which compound is a strong acid?
   a. $HClO_2(aq)$
   b. $H_2CO_3(aq)$
   c. formic acid
   d. perchloric acid
5. Which one of the following compounds is not a strong electrolyte?
   a. $CH_3COOH(aq)$
   b. $HClO_4(aq)$
   c. $HI(aq)$
   d. $NaOH(aq)$
6. Which of the following is usually referred to as strong acid in water solution? Write the ionization reactions.
   a. HF
   b. HNO_2
   c. H_2CO_3
   d. HSO_4^-
   e. HNO_3
   f. $HClO_4$

25.3 Arrhenius Bases

Key Concepts

In this lesson students explore the Arrhenius base definition and properties of bases.

Lesson Objectives

- Define an Arrhenius base and know some examples of bases.
- State the properties of bases.
- Describe the neutralization reaction that bases undergo.

Lesson Vocabulary

Arrhenius base A substance that produces OH^- ions in a solution.

Strategies to Engage

- Before beginning this lesson, review with students the Arrhenius definition of an acid. Have students recall the formula for most bases and see if they can come up with the Arrhenius definition of a base. Tell students to "stay tuned" to see if their definition was correct.

Strategies to Explore

- After exploring how bases affect indicators, have groups of students write the procedure to perform a color-change trick using acids, bases, and indicators. After their procedure is approved, have the students use their procedure to perform a chemistry magic show.

Strategies to Extend and Evaluate

- Have students do library research on the topic of acids and bases in photography and prepare a written report, Keynote or PowerPoint slideshow, or display.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 25.3 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. What is the role of litmus paper for acids and base chemistry?
2. What are the properties of bases? Give a common example.
3. Which statement best describes a characteristic of a base solutions?
   a. They taste bitter.
   b. They turn red litmus paper red.
   c. They react with some metals to form hydrogen gas.
   d. They are weak electrolytes.
4. Which of the following is the Arrhenius definition of a base?
   a. A base is a substance that donates protons.
   b. A base is a substance that accepts protons.
   c. A base is a substance that dissolves in water to form OH^- ions.
   d. A base is a substance that reacts with water to form H^+ ions.
5. Which of the following bases would be a weak electrolyte?
   a. $NaOH$
   b. $Ba(OH)_2$
   c. $Ca(OH)_2$
   d. $Al(OH)_3$
6. Write the balanced neutralization reaction between the following acids and bases.
   a. potassium hydroxide + hypochlorous acid
   b. hydrobromic acid + calcium hydroxide
   c. hydrochloric acid + sodium hydroxide
   d. potassium hydroxide + sulfuric acid
7. Write the net ionic equation for each of the neutralizations reactions in #4.

25.4 Salts

Key Concepts

In this lesson students explore acid-base neutralization reactions.

Lesson Objectives

- Describe the formation of salts in neutralization reactions in terms of Arrhenius theory.
- Identify acidic, basic, and neutral salts from neutralization reaction.

Lesson Vocabulary

basic salt A salt formed in a neutralization reaction between a weak acid and a strong base.

acidic salt A salt formed in a neutralization reaction between a strong acid and a weak base.

neutral salt A salt formed in a neutralization reaction between a strong acid and a strong base or a weak acid and a weak base.

Strategies to Engage

Strategies to Explore

- Have students create a flow chart that can be used to determine the type of salt that will form from an acid/base neutralization reaction. Encourage interested students to show their flow charts to the class and have the class choose the best one.

Strategies to Extend and Evaluate

- Have students work in pairs or teams to write a poem about acids, bases, and salts. Their poems should explain what they are, some of their properties, and how they differ from each other.
- Have students do library research on the topic of natural solutions to acid rain and prepare a written report, Keynote or PowerPoint slideshow, or display.
- Have students complete the lab ***Hydrolysis of Salts***. This lab is located in the Supplemental Lab Book.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 24.4 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. How do an acid and a base fit the definition of an ionic compound? Use examples in your answer.
2. Explain neutralization reactions in terms of Arrhenius theory. Use an example in your answer.
3. Which salt will form a basic solution when dissolved in water?
    a. KNO_3
    b. $CaCl_2$
    c. $NaClO_4$
    d. $NaNO_2$
4. Which salt will form an acidic solution when dissolved in water?
    a. copper(II) sulfate
    b. sodium acetate
    c. potassium chloride
    d. sodium cyanide
5. Milk of magnesia is a common over- the - counter antacid that has, as its main ingredient, magnesium hydroxide. It is used by the public to relieve acid indigestion. Acid indigestion is caused by excess stomach acid being present. Since a stomach upset is caused by excess hydrochloric acid, this tends to be a quite painful affliction for people. Write the balanced chemical equation for the reaction between milk of magnesia and hydrochloric acid. What type of reaction is this? What type of salt is formed?
6. Complete the following neutralization reactions and identify the type of salt produced.
    a. $H_2SO_4(aq)$ and $NaOH(aq) \rightarrow$
    b. $HNO_3(aq)$ and $NH_4OH(aq) \rightarrow$
    c. $HF(aq)$ and $NH_4OH(aq) \rightarrow$
    d. $CH_3COOH(aq)$ and $KOH(aq) \rightarrow$
    e. $HCl(aq)$ and $KOH(aq) \rightarrow$

25.5 pH

Key Concepts

In this lesson students explore $[H^+]$ and $[OH^-]$ and pH.

Lesson Objectives

- Calculate $[H^+]$ for strong acids and $[OH^-]$ for strong bases.
- Define autoionization and use it to find $[H^+]$ from $[OH^-]$ or to find $[OH^-]$ from $[H^+]$.
- Describe the pH scale.
- Define pH.
- Calculate pH from $[H^+]$ or vice versa.

Lesson Vocabulary

pH scale A scale measuring the $[H^+]$ with values from 0 to 14.

$pH = -\log [H^+]$ - Formula used to calculate the power of the hydronium ion.

autoionization When the same reactant acts as both the acid and the base.

ion product constant for water K_w, is the product of the hydronium ion and the hydroxide ion concentrations in the autoionization of water.

Strategies to Engage

- Students are likely to have heard about pH in advertising and popular media (e.g., shampoos, antacids). Call on volunteers to share with the class anything they already know about pH. Point out correct responses and clear up any misconceptions. Tell students they will learn more about pH in this lesson.

Strategies to Explore

- Review acidic, basic, and neutral salts. Use a pH meter to demonstrate pH of salts of weak acids and bases.

Strategies to Extend and Evaluate

- Ask students to come up with their own lesson review questions. Then have them exchange papers with a classmate to have students answer each others questions.
- Have students write a paragraph about a person preparing meals in a restaurant. In their paragraphs, ask them to describe various solutions as acids or bases, and estimate the pH values.
- Encourage interested students to do library research on Soren Sorensen and the concept of pH. Students should be prepared to share their findings with the class.
- Encourage interested students to do library research on the topic of pH and home food canning. Students should be prepared to share their findings with the class.
- Have groups of students research and prepare natural acid-base indicators such as those that can be made from cabbage juice, cherries and tumeric. Students should perform an in-class demonstration of their chosen indicator.

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled pH. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 25.5 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

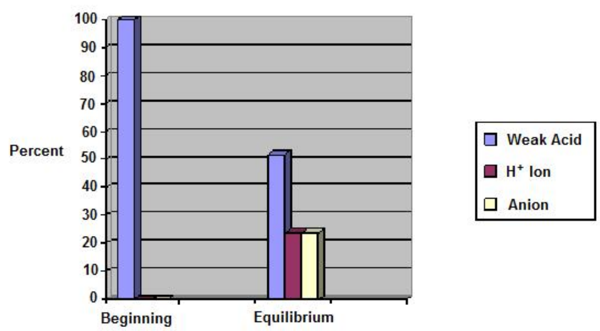
1. Why is it necessary to balance the chemical equation before determining the $[H^+]$ or $[OH^-]$ for strong acids and bases?
2. Why can't you determine the $[H^+]$ or $[OH^-]$ for weak acids and bases the same way you can determine the $[H^+]$ or $[OH^-]$ for strong acids and bases?
3. What is the $[H^+]$ ion concentration in a solution of 0.350 mol/L H_2SO_4?
   a. 0.175 mol/L
   b. 0.350 mol/L
   c. 0.700 mol/L
   d. 1.42×10^{-14} mol/L
4. A solution has a pH of 6.54. What is the concentration of hydronium ions in the solution?
   a. 2.88×10^{-7} mol/L
   b. 3.46×10^{-8} mol/L
   c. 6.54 mol/L
   d. 7.46 mol/L
5. A solution has a pH of 3.34. What is the concentration of hydroxide ions in the solution?
   a. 4.57×10^{-4} mol/L
   b. 2.19×10^{-11} mol/L
   c. 3.34 mol/L
   d. 10.66 mol/L

6. A solution contains 4.33×10^{-8} M hydroxide ions. What is the pH of the solution?
   a. 4.33
   b. 6.64
   c. 7.36
   d. 9.67
7. Fill in the **Table** 25.4 and rank the solutions in terms of increasing acidity.

TABLE 25.4:

Solutions	$[H^+]$ (mol/L)	$-\log\ [H^+]$	pH
A	0.25	0.60	0.60
B	?	2.90	?
C	1.25×10^{-8}	?	?
D	0.45×10^{-3}	?	?
E	?	1.26	?

8. A bottle of calcium hydroxide is found in the lab with a label reading: 0.014 mol/L.
   a. What are the concentrations of all of the ions present in the solution?
   b. What is the pH of the solution?
9. It has long been advocated that red wine is good for the heart. However, wine is also considered to be an acidic compound. Determine the concentration of hydronium ions in wine with pH3.81.
10. The diagram that follows represents a weak acid before ionization and when the reaction comes to equilibrium. If the acid were weaker than the one represented in the diagram, how would the diagram change? Draw a new diagram to represent your answer. If the acid were a strong acid, how would the diagram change? Draw a new diagram to represent your answer.



25.6 Weak Acid/Base Equilibria

Key Concepts

In this lesson students explore weak acids and weak bases as equilibrium systems.

Lesson Objectives

- Define weak acids and weak bases in terms of equilibrium.
- Define K_a and K_b.
- Use K_a and K_b to determine acid and base strength.
- Use K_a and K_b in acid/base equilibrium problems.

Lesson Vocabulary

acid ionization constant K_a represents the equilibrium constant for the ionization of a weak acid.

base dissociation constant K_b represents the equilibrium constant for the dissociation of a weak base.

Strategies to Engage

- Before beginning this lesson, review with students equilibrium constants. Explain to students that in this lesson, they will explore equilibrium constants for acids and bases. Give students the opportunity to define the vocabulary terms just by examining the words themselves. At the end of the lesson, have students check their original definition and make corrections, if necessary.

Strategies to Explore

- Use the mathematical calculations in this lesson to reduce the reliance on language skills. As you go through each example problem, use them to explain the concepts explored in this lesson. DI **English Language Learners**

Strategies to Extend and Evaluate

- On the board or chart paper, have students write a class summary of this lesson. Have one student come up with the first sentence, and have students contribute sentences until the entire lesson has been summarized.

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Weak Acids and Bases**. Ask students to complete the worksheet alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 25.6 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. What makes weak acids and bases a special case for equilibrium reactions?
2. What do the constants K_a and K_b represent?
3. Oxalic acid is a weak acid. Its ionization reaction is represented below. $H_2C_2O_{4(aq)} + H_2O_{(L)} \leftrightarrows H_3O^+{}_{(aq)} + HC_2O_4{}^-{}_{(aq)}$ Which of the following best represents the acid ionization constant expression, K_a?
   a. $K_a = \frac{[H_3O^+][HC_2O_4^-]}{[H_2C_2O_4][H_2O]}$
   b. $K_a = \frac{[H_2C_2O_4][H_2O]}{[H_3O^+][HC_2O_4^-]}$
   c. $K_a = \frac{[H_3O^+][HC_2O_4^-]}{[H_2C_2O_4]}$
   d. $K_a = \frac{[H_2C_2O_4]}{[H_3O^+][HC_2O_4^-]}$
4. Choose the weakest acid from the list below.
   a. $HNO_{2(aq)}; K_a = 5.6 \times 10^{-3}$
   b. $HF_{(aq)}; K_a = 6.6 \times 10^{-4}$
   c. $H_3PO_{4(aq)}; K_a = 6.9 \times 10^{-3}$
   d. $HCOOH_{(aq)}; K_a = 1.8 \times 10^{-4}$
5. Choose one of the following reactions that would best represent a reaction that has an equilibrium constant best described as a base dissociation constant, K_b.
   a. $H_2PO_4{}^-{}_{(aq)} + H_2O_{(L)} \leftrightarrows HPO_4{}^{-2}{}_{(aq)} + H_3O^+{}_{(aq)}$
   b. $NH_4{}^+{}_{(aq)} + H_2O_{(L)} \leftrightarrows NH_{3(aq)} + H_3O^+{}_{(aq)}$
   c. $NH_4{}^-{}_{(aq)} + OH^-{}_{(aq)} \leftrightarrows NH_{3(aq)} + H_2O_{(L)}$
   d. $F^-{}_{(aq)} + H_2O_{(L)} \leftrightarrows HF_{(aq)} + OH^-{}_{(aq)}$
6. A 0.150 mol/L solution of a weak acid having the general formula HA is 15.0% ionized in aqueous solution. Which expression best represents the calculation of the acid ionization constant K_a for this acid?
   a. $K_a = \frac{(0.150)(0.150)}{(0.150)}$
   b. $K_a = \frac{(0.0225)(0.0225)}{(0.128)}$
   c. Not enough information is given.

7. Put the following bases in order of increasing base strength. Write equilibrium reactions for each. ethanolamine $(HOCH_2CH_2NH_2), K_b = 3.2 \times 10^{-5}$, piperidine $(C_5H_{10}NH), K_b = 1.3 \times 10^{-3}$, triethylamine $((CH_3CH_2)_3N), K_b = 5.2 \times 10^{-4}$, and ethylenediamine $(H_2NCH_2CH_2NH_2), K_b = 8.5 \times 10^{-5}$.

25.7 Bronsted Lowry Acids-Bases

Key Concepts

In this lesson students explore the Bronsted-Lowry definitions of acids and bases and acid-base conjugate pairs.

Lesson Objectives

- Define a Brønsted-Lowry acid and base.
- Identify Brønsted-Lowry acids and bases from balanced chemical equations.
- Define conjugate acid and conjugate base.
- Identify conjugate acids-bases in balanced chemical equations.
- Identify the strength of the conjugate acids and bases from strengths of the acids and bases.

Lesson Vocabulary

Brønsted-Lowry acid A substance that donates a proton (H^+).

Brønsted-Lowry base A substance that accepts a proton (H^+).

amphoteric substances Substances that act as both acids and bases in reactions (i.e. NH_3).

conjugate acid The substance that results when a base gains (or accepts) a proton.

conjugate base The substance that results when an acid loses (or donates) a proton.

Strategies to Engage

- Review with students the Arrhenius acid/base definitions. Explain to students that in this lesson they will explore a more generalized definition. Encourage students to give reasons why a more generalized definition was necessary.

Strategies to Explore

Strategies to Extend and Evaluate

- As a review of this lesson, have students record what they think is the main idea of each section. Have pairs of students come to a consensus on each main idea. Then, have each pair combine with another pair and again come to a consensus. Finally, have each group share their results with the class. DI **Less Proficient Readers**
- Have students create a concept map showing the relationships among the lesson vocabulary terms.

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Conjugate Acids-Bases**. Ask students to complete the worksheet alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 25.7 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. What improvements did Brønsted-Lowry make over the Arrhenius definition for acids-bases?
2. If you were to use $HCl(aq)$ as an example, how would you compare the Arrhenius definition of an acid to the Brønsted-Lowry definition?
3. What is the Brønsted-Lowry definition of an acid?
   a. a substance that donates protons
   b. a substance that accepts protons
   c. a substance that dissolves in water to form OH^- ions
   d. a substance that dissolves in water to form H^+ ions
4. If H_3O^+ is an acid according to the Brønsted-Lowry theory, what is the conjugate base of this acid?
   a. $H_4O^{2+}{}_{(aq)}$
   b. $H^+{}_{(aq)}$
   c. $H_2O_{(L)}$
   d. $OH^-{}_{(aq)}$
5. What is the conjugate base of $H_2PO_4{}^-$?
   a. $H_3O^+{}_{(aq)}$
   b. $H_3PO4_{(aq)}$
   c. $HPO_4{}^{2-}{}_{(aq)}$
   d. $PO_4{}^{3-}{}_{(aq)}$
6. In the following reactions, which are the Brønsted-Lowry acids? i) $H_3PO_{4(aq)} + H_2O_{(L)} \leftrightarrows H_3O^+{}_{(aq)} + H_2PO_4{}^-{}_{(aq)}$ ii) $H_2PO_4{}^-{}_{(aq)} + H_2O_{(L)} \leftrightarrows H_3O^+{}_{(aq)} + HPO_4{}^{2-}{}_{(aq)}$
   a. $H_2PO_4{}^-, H_2O, HPO_4{}^{2-}$

b. $H_3PO^4, H_2O, H_2PO_4{}^-$
c. $H_3O^+, H_2O, HPO_4{}^{2-}$
d. $H_3PO4, H_3O^+, H_2PO_4{}^-$

7. Label the conjugate acid-base pairs in each reaction.
   a. $HCO_3{}^- + H_2O \leftrightarrows H_2CO_3 + OH^-$
   b. $H_2PO_4{}^- + H_2O \leftrightarrows H_3O^+ + HPO_4{}^{2-}$
   c. $CN^- + H_2O \leftrightarrows HCN + OH^-$
   d. $HF(aq) + H_2O(l)\ \Omega\ H_3O^+(aq) + F^-(aq)$
8. Complete the following reactions. When done, label the conjugate acid/base pairs.
   a. $BrO_3{}^- + H_2O \leftrightarrows$
   b. $HNO_3 + H_2O \rightarrow$
   c. $HSO_4{}^- + C_2O_4{}^{2-} \leftrightarrows$
9. For the reactions in question 7, which are the weak conjugate bases and which are the strong conjugate bases?

25.8 Lewis Acids and Bases

Key Concepts

In this lesson students explore the Lewis definitions of acids and bases, and reactions of Lewis acids and bases.

Lesson Objectives

- Define a Lewis acid and a Lewis base.
- Define a coordinate covalent bond.
- Identify a Lewis acid and a base in reactions.

Lesson Vocabulary

Lewis acid A substance that accepts a pair of electrons from a substance (i.e. BF_3).

Lewis base A substance that donates a pair of electrons from a substance (i.e. NH_3).

coordinate covalent bond A covalent bond formed where both electrons that are being shared come from the same atom.

Strategies to Engage

- Have students recall what they know about Gilbert Lewis. They may recall that he is the scientist after whom Lewis structures are named. Review Lewis structures with students and have students try to figure out what the Lewis acid-base definitions focus on.

Strategies to Explore

Strategies to Extend and Evaluate

- Have students use grid paper to make a crossword puzzle using the vocabulary terms. Ask students to exchange papers with a classmate and solve each other's puzzles.
- Have students write a paragraph explaining the three acid-base theories explored in this chapter.

- Ask students to search for examples of bad science about acids and bases on the web or in books. Have them quote the claim, reference the source, and then explain what is wrong.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 25.8 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. How do the Lewis definitions of acids and bases compare to the Brønsted-Lowry definitions of acids and bases?
2. In the following reversible reaction, which of the reactants is acting as a Lewis base? $Cd^{2+}{}_{(aq)} + 4I^{-}{}_{(aq)} \leftrightarrows CdI_4{}^{2-}{}_{(aq)}$
   a. Cd^{2+}
   b. I^{-}
   c. $CdI_4{}^{2-}$
   d. none of the above, this is not an acid-base reaction.
3. Which of the following statements are false?
   a. NH_3 is a Lewis base.
   b. $B(OH)_3$ is a Lewis acid.
   c. CO_2 is a Lewis base.
   d. Ag^{+} Is a Lewis acid.
4. Which of the following statements are true?
   a. NH_3 is a Lewis base.
   b. $B(OH)_3$ is a Lewis acid.
   c. CO_2 is a Lewis base.
   d. Ag^{+} Is a Lewis acid.
5. Classify each of the following as a Lewis acid or base.
   a. H_2O
   b. BF_3
   c. S^{2-}
   d. Cu^{2+}
   e. O^{2-}
6. Write the balanced chemical equation between $SO_3{}^{2-}$ and H_2O and label the Lewis acids and bases.
7. Identify the Lewis acid and Lewis base in each of the following reactions. Then write the reactions.
   a. $Cu^{2+}{}_{(aq)} + 6\,H_2O_{(L)} \rightarrow Cu(H_2O)_6{}^{2+}{}_{(aq)}$
   b. $(CH_3CH_2)_2O_{(aq)} + AlCl_{3(aq)} \rightarrow (CH_3CH_2)_2OAlCl_{3(aq)}$

Chapter 26 Water, pH, and Titration TE

Chapter Outline

Water, pH, and Titration

Outline

This chapter *Water, pH, and Titration* consists of four lessons that cover the mathematics of the dissociation of water, acid-base indicators, acid-base titration, and buffers.

- **Lesson 26.1 Water Ionizes**
- **Lesson 26.2 Indicators**
- **Lesson 26.3 Titration**
- **Lesson 26.4 Buffers**

Overview

In these lessons, students will explore:

- The autoionization of water, and the mathematics of pH and pOH.
- Natural and synthetic indicators.
- How the process of titration is used to determine the concentration of acids and bases.
- The chemistry of buffer solutions.

Science Background Information

Acid Rain

The exceptional characteristics of the substance known as water have been recognized and appreciated for millennia. In particular, its ability as a solvent provides for many of the vital processes enabling life, such as acting as the medium in which red blood cells transport oxygen in our bodies. Yet water's propensity to dissolve ions, other liquids and even gases may not always produce physical or biochemical advantages.

The water supply on Earth is continuously transported and concomitantly purified by a mechanism known as the hydrological cycle. As solar radiation heats the Earth's surface, water molecules evaporate and then condense

into cloud formations as they reach higher elevations and cooler atmospheric levels. Large-scale weather patterns transport the water in these cloud formations around the globe, and return the water to the surface as precipitation. Despite the purification of the substance by this process, rainwater is found to have a pH that is not neutral as one might expect, but mildly acidic. During its passage through the atmosphere, water's extraordinary capacity as a solvent absorbs carbon dioxide in the air, and small quantities of carbonic acid is generated as shown:

$$H_2O + CO_2 \rightarrow H_2CO_3 \rightarrow H^+ + HCO_3^-$$

In our modern industrialized world, there are other gases present in the atmosphere that, like CO_2, can dissolve in atmospheric moisture. In particular, the presence of NO_x and SO_x, byproducts of fossil fuel combustion, is a specific concern.

NO_x, formed by the reaction of nitrogenous contaminants in fuels with oxygen, can react with water in the atmosphere to generate nitric acid, HNO_3. In its concentrated form, nitric acid is a corrosive material that can dissolve some metals. Likewise, sulfur oxide contaminants react with moisture yielding sulfuric acid, the viscous acid found in lead-acid car batteries. As these acids are produced and dispersed in the atmosphere, they constitute an environmental issue that transcends borders and physical boundaries.

Acid rain is then precipitation that possesses acidity greater than that of normally slightly acidic rainwater. Its effects can be noted on both biological systems and physical structures. Trees in many areas of the world bear the evidence of acid rain damage in the form of brittle, browned leaves, but the principal destruction to plants is to the root system. Increased acidity limits access of beneficial ions such as Ca^{2+} and Mg^{2+} but encourages the solubility of damaging ions such as Al^{3+} in the soil. Aquatic organisms such as fish experience skeletal growth problems due to limited access to calcium ions. Physical structures also bear witness to the destructive nature of acid rain; marble statues erode due to long–term exposure to acidic rainfall.

Pacing the Lessons

Use the table below as a guide for the time required to teach the lessons of *Water, pH, and Titration*.

TABLE 26.1: 60 Minute Class Periods per Lesson

Lesson	Number of Class Periods
26.1 *Water Ionizes*	0.5
26.2 *Indicators*	1.0
26.3 *Titration*	2.0
26.4 *Buffers*	1.5

Managing Materials

The following items are needed to teach the strategies and activities described in the Teachers Edition of the FlexBook for *Water, pH, and Titration*.

TABLE 26.2: Water, pH, and Titration Materials List

Lesson	Strategy or Activity	Materials Needed
26.1		
26.2		
26.3		
26.4		

Multimedia Resources

You may find these additional internet resources helpful when teaching *Water, pH, and Titration*:

- Lesson on buffer solutions: http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/buffer12.swf
- List of household acid-base indicators: http://antoine.frostburg.edu/chem/senese/101/acidbase/faq/household-indicators.shtml
- pH calculation problem generator: http://science.widener.edu/svb/tutorial/phcalcs.html
- Autoionization of water animation: http://chemmovies.unl.edu/ChemAnime/AUTOWD/AUTOWD.html

Possible Misconceptions

Making the FlexBook Flexible

An important advantage of the FlexBook is the ability it gives you, the teacher, to select the chapters and lessons that you think are most important for your own classes. You should also consult the standards correlation table that follows when selecting chapters and lessons to include in the FlexBook for your classes.

TABLE 26.3: Standard Addressed by the Lessons in Water, pH, and Titration

Lesson	California Standards	NSES Standards	AAAS Benchmarks
26.1			
26.2	5d		
26.3	1b, 5d		
26.4	5g		

26.1 Water Ionizes

Key Concepts

In this lesson students explore the autoionization of water and the mathematics of pH and pOH.

Lesson Objectives

- The students will write the equation for the autoionization of water, and express the concentration of hydrogen and hydroxide ion in a neutral solution at $25^\circ C$.
- The students will express the value of K_w in a water solution at $25^\circ C$.
- The students will write the formulas for pH and pOH, and show the relationship between these values and K_w.
- The students will express the relationship that exists between pH, pOH, and K_w.
- Given the value of any one of the following values in a water solution at $25^\circ C$, the students will calculate all the other values; $[H^+]$, $[OH^-]$, pH, and pOH.
- The students will state the range of values for pH that indicate a water solution at $25^\circ C$ is acidic.
- The students will state the range of values for pH that indicate a water solution at $25^\circ C$ is basic.
- The students will state the range of values for pH that indicate a water solution at $25^\circ C$ is neutral.

Lesson Vocabulary

autoionization Autoionization is when the same reactant acts as both the acid and the base.

Strategies to Engage

- Review the Bronsted-Lowry definition of acids and bases. Write the formula for water on the board in the form of HOH. Have students attempt to use these two concepts to write an equation showing the autoionization of water.

Strategies to Explore

Strategies to Extend and Evaluate

- Challenge interested students to come up with their own questions addressing this objective; "Given the value of any one of the following values in a water solution at $25°C$, the student will calculate all the other values; $[H^+]$, $[OH^-]$, pH, and pOH. Students may then exchange papers with a classmate to have them answer each other's questions.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 26.1 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. How do pH and pOH relate to the pH scale?
2. What does the value of K_w tell you about the autoionization of water?
3. If the pH of an unknown solution is 4.25, what is the pOH?
   a. $10^{-4.25}$
   b. $10^{-9.75}$
   c. 9.75
   d. $14.0 - 10^{-9.75}$
4. A solution contains a hydronium ion concentration of 3.36×10^-4 mol/L. What is the pH of the solution?
   a. 3.36
   b. 3.47
   c. 10.53
   d. none of the above
5. A solution contains a hydroxide ion concentration of 6.43×10^{-9} mol/L. What is the pH of the solution?
   a. 5.80
   b. 6.48
   c. 7.52
   d. 8.19
6. An unknown solution was found in the lab. The pH of the solution was tested and found to be 3.98. What is the concentration of hydroxide ion in this solution?

a. 3.98 mol/L
b. 0.67 mol/L
c. 1.05×10^{-4} mol/L
d. 9.55×10^{-11} mol/L

7. If a solution is known to have a hydroxide ion concentration of 2.5×10^{-5} mol/L, then the pH of the solution is ___________ and it is ___________.

a. 2.5, acidic
b. 4.6, basic
c. 4.6, acidic
d. 9.4, basic

8. K_w is the ionization product constant for water but is also the equilibrium constant for the acid-base autoionization reaction for water. When dealing with equilibrium constants, such as K_w, it is important to take into account the temperature as temperature affects the value of the equilibrium constant. The value of 1.0×10^{-14} for K_w is for a temperature of $25°C$. If the temperature was raised to $60°C$, the value of K_w changes to 1.0×10^{-13}. How does this effect $[H^+]$,$[OH^-]$, pH, and pOH?

26.2 Indicators

Key Concepts

In this lesson students explore natural and synthetic indicators.

Lesson Objectives

- Define an acid-base indicator.
- Explain the difference between natural and synthetic indicators.
- List examples of natural and synthetic indicators.
- Explain how indicators work.
- Explain the usefulness of indicators in the lab.

Lesson Vocabulary

indicator A substance that changes color at a specific *pH* and is used to indicate the *pH* of the solution.

natural indicator An indicator that is produced from a substance that is naturally occurring, or is itself a naturally occurring substance.

synthetic indicator An indicator that is a complicated structure of an organic weak acid or base.

Strategies to Engage

- Have students read the review questions at the end of this section. This way, students will be familiar with the types of information that they will explore in this section.

Strategies to Explore

- Have students write a short paragraph comparing and contrasting natural and synthetic indicators. They should briefly explain the properties for each type of indicator.

Strategies to Extend and Evaluate

- Have students work in pairs or teams to write a poem about indicators. Their poems should explain what indicators are, how they work, and give some examples of natural and synthetic indicators.
- Outline the main concepts of the lesson as a class. Discuss the main concepts as you prepare the outline.
- Have students complete the lab *pH**Measurement Using Indicators***. This lab is located in the Supplemental Lab Book.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 26.2 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Describe the uses for litmus and universal indicator in the laboratory setting.
2. What is the difference between a natural and a synthetic indicator?
3. Describe how indicators work.
4. If you had an acid-base neutralization reaction that turned phenolphthalein pink and Thymolphthalein blue, what is the pH of the solution?
   a. 8.2
   b. 9.4
   c. 10
   d. Not enough information is available.
5. If you had an acid-base neutralization reaction that turned methyl violet blue and Thymol blue orange, what is the pH of the solution?
   a. 1.6
   b. 2.0
   c. 2.8
   d. Not enough information is available.
6. Universal indicator is an indicator commonly used in the laboratory. At a pH of 6 it is pale yellow and at a pH of 4 it is pale orange. If the indicator was orange, which statement would be definitely true?
   a. The solution is probably acidic.
   b. The pH is between 4 and 5.
   c. The solution is probably basic.
   d. The pH is less than 5.0.
7. Alizarin Yellow R is an indicator that changes color in the pH range from 10.1 to 12.0. Below 10.1 the color is Yellow, above 12.0 the color is red. If the color of the solution containing Alizarin Yellow R was orange, which statement about the solution would be true?
   a. The pH is below 10.

b. The *pH* is above 12.0.
c. The solution is definitely acidic.
d. The *pH* is between 10.1 and 12.0.

8. If the *pH* of the solution is 8.9, what would be the color of the solution if the following indicators were added?
    a. Universal indicator
    b. Thymol blue
    c. Methyl blue
9. A solution has been found in the laboratory that was tested with a number of indicators. It was found that the following indicators showed these results:
    a. Phenolphthalein was colorless
    b. Orange IV was yellow
    c. Universal indicator was orange
    d. Methyl orange was red

    What was the *pH* of the solution?

26.3 Titrations

Key Concepts

In this lesson students explore how the process of titration is used to determine the concentration of acids and bases.

Lesson Objectives

- Define titrations and identify the different parts of the titration process.
- Explain the difference between the endpoint and the equivalence point.
- Describe the three types of titration curves.
- Identify points on the titration curves for the three types of titrations.
- Define a standard solution.
- Calculate the accurate concentration of an acid or base using a standard.
- Calculate unknown concentrations or volumes of acids or bases at equivalence.

Lesson Vocabulary

titration The lab process in which a known concentration of base (or acid) is added to a solution of acid (or base) of unknown concentration.

titrant The solution in the titration of known concentration.

burette A piece of equipment used in titrations to accurately dispense the volume of the solution of known concentration (either a base or an acid).

Erlenmeyer flask A piece of equipment used in titrations (and other experiments) to hold a known volume of the unknown concentration of the other solution (either the acid or the base).

endpoint The point in the titration where the indicator changes color.

equivalence point The point in the titration where the number of moles of acid equals the number of moles of base.

A device used to measure the changes in pH as the titration goes from start to finish.

titration curve A graph of the pH versus the volume of titrant added.

standard solution A solution whose concentration is known exactly and is used to find the exact concentration of the titrant.

Strategies to Engage

- Review some of the prior knowledge students have obtained about acids and bases, about chemical reactions, molarity calculations, and about indicators that apply to the concept of titrations.

Strategies to Explore

- Challenge interested students to describe the process of titration as concisely and correctly as possible. Have the rest of the class choose the student who is able to correctly describe the process using the least amount of words.
- Have students complete the lab ***Acid-Base Titration***. This lab is located in the Supplemental Lab Book.

Strategies to Extend and Evaluate

- Have students use grid paper to make a crossword puzzle using the vocabulary terms. Ask students to exchange papers with a classmate and solve each other's puzzles.
- Have students write a one-paragraph summary of this lesson. Instruct students to correctly use each vocabulary term in their summary.
- Have students use titration to test antacids quantitatively. Divide students into groups and ask them to do research to find a suitable experimental procedure. After their procedures have been approved, allow them to perform their procedures in class.

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Titration**. Ask students to complete the worksheet alone or in pairs as a review of lesson content.

Review Questions

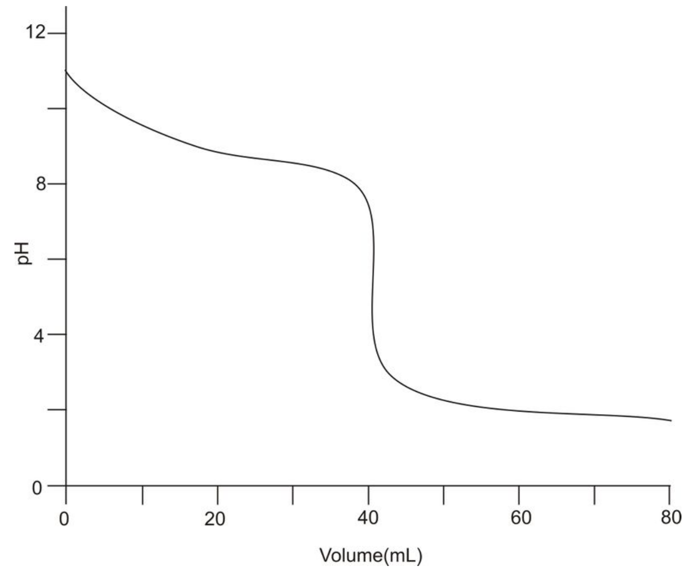
Have students answer the Lesson 26.3 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Why do you think there would be more experimental error with using an indicator over using a *pH* meter in a titration?
2. Why would there not be a weak acid-weak base titration?
3. Which of the following definitions best suits that of an endpoint?
   a. The stoichiometric point where the number of moles of acid equals the number of moles of base.
   b. The visual stoichiometric point where the number of moles of acid equals the number of moles of base.
   c. The midpoint of the vertical stretch on the titration curve.
   d. None of the above

4. In the following titration curve, what pair of aqueous solutions would best represent what is shown to be happening in the curve?
   a. $HCOOH_{(aq)} + NH_{3(aq)}$
   b. $HCOOH_{(aq)} + NaOH_{(aq)}$
   c. $H_2SO_{4(aq)} + Ba(OH)_{2(aq)}$
   d. $HClO_{4(aq)} + NH_{3(aq)}$



5. What would be the best indicator to choose for the pH curve shown in question 3?
   a. Methyl red
   b. Litmus
   c. Phenolphthalein
   d. Phenol red
6. What is the best indicator to use in the titration of benzoic acid with barium hydroxide?
   a. Methyl violet
   b. Bromothymol blue
   c. Phenolphthalein
   d. Methyl blue
   e. Indigo carmine

TABLE 26.4:

Indicator	pH Range
Methyl violet	0.0 – 1.6
Bromothymol blue	3.0 – 4.7
Phenolphthalein	8.2 – 10.0
Methyl blue	10.6 – 13.4
Indigo carmine	11.4 – 13.0

7. If 22.50 mL of a sodium hydroxide is necessary to neutralize 18.50 mL of a 0.1430 mol/LHNO_3 solution, what is the concentration of $NaOH$?
   a. 0.1176 mol/L
   b. 0.1430 mol/L
   c. 0.1740 mol/L
   d. 2.64 mol/L

8. Plot the following titration data on a titration curve of pH vs. volume of base added. When complete, find the pH at equivalence and choose an appropriate indicator for the titration. What volume of base is necessary to neutralize all of the acid?

TABLE 26.5:

Volume of base added (mL)	pH
0.00	1.0
2.00	1.2
4.00	1.4
6.00	1.6
8.00	1.9
9.00	2.3
9.50	2.6
9.90	3.3
9.99	4.3
10.00	7.0
10.01	9.7
10.50	10.7
12.00	11.4
14.00	12.1
16.00	12.1

9. Calculate the concentration of hypochlorous acid if 25.00 mL of $HClO$ is used in a titration with 32.34 mL of a 0.1320 mol/L solution of sodium hydroxide.

26.4 Buffers

Key Concepts

In this lesson students explore the chemistry of buffer solutions.

Lesson Objectives

- Define a buffer and give various examples of buffers.
- Explain the effect of a strong acid on the pH of a weak acid/conjugate base buffer.
- Explain the effect of a strong base on the pH of a weak base/conjugate acid buffer.

Lesson Vocabulary

buffer A buffer is a solution of a weak acid and its conjugate base or a weak base and its conjugate acid that resists changes in pH when an acid or base is added to it.

Strategies to Engage

Strategies to Explore

- Have students write down the lesson objectives, leaving about 5 or 6 lines of space in between. As you explore the lesson, have students write the "answer" to each objective.

Strategies to Extend and Evaluate

- Have students work in pairs to create an advertisement for a "buffered" aspirin. It should resemble an ad that might appear in a newspaper or a magazine. Students should illustrate their ad and write a slogan to explain why the "buffered" aspirin is preferred over regular aspirin.
- Have students write a paper describing how buffers are an application of Le Chatêlier's Principle.
- Have students work in pairs or teams to write a poem about buffers. Their poems should explain what they are, how they work, and give some examples.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 26.4 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Define a buffer solution.
2. What are two different types of buffer solutions?
3. One of the following statements of buffers is incorrect. Which one?
   a. A buffer may be prepared from a weak acid and its conjugate base salt.
   b. A buffer may be prepared from a weak base and its conjugate acid salt.
   c. A buffer is a solution that can resist changes in pH when any amount of acid or base is added to it.
   d. A buffer is a solution that can resist changes in pH when a small amount of acid or base is added to it.
4. Which pair of aqueous 1.0 mol/L solutions could be chosen to prepare a buffer?
   1. i and iii only
   2. i and iii only
   3. i, ii and iii
   4. None of these solutions is a buffer.
      a. $NH_4HSO_{4(aq)}$ and $H_2SO_{4(aq)}$
      b. $HNO_{2(aq)}$ and $NaNO_{2(aq)}$
      c. $NH_4Cl_{(aq)}$ and $NH_{3(aq)}$
5. Which of the following would form a buffer solution if combined in appropriate amounts?
   a. HCl and $NaCl$
   b. HCN and $NaCN$
   c. H_2S and Na_2S
   d. HNO_3 and $NaNO_3$
6. A buffer is made up of a weak acid and a conjugate base. A small amount of acid is added to the buffer. What happens to the resulting solution?
   a. The acid dissociation constant goes up.
   b. The concentration of the weak acid in the buffer goes down.
   c. The pH of the solution goes up.
   d. The pH remains almost the same.
7. Almonds from the wild have a very bitter taste because of hydrogen cyanide (and therefore are very dangerous to eat!!!). Interestingly if we think about HCN in a buffer situation, HCN and $NaCN$ can be considered to act as a buffer solution. Sulfurous acid is used quite frequently as a cleansing agent. If we take the sodium salt, Na_2SO_3, of sulfurous acid, H_2SO_3 we do not make a buffer solution. Why is this so? Why would one make a buffer solution and not the other?

CHAPTER 27 Thermodynamics - HS Chemistry TE

Chapter Outline

Unit 10 Thermodynamics

Outline

This unit, Thermodynamics, includes one chapter that covers the energy involved in bond breaking and bond formation, the heat of reaction, the heat of formation, Hess's law, entropy, and Gibb's free energy.

- **Chapter 27: Thermodynamics**

Overview

Thermodynamics

This chapter covers the energy involved in bond breaking and bond formation, the heat of reaction, the heat of formation, Hess's law, entropy, and Gibb's free energy.

Thermodynamics

Outline

This chapter *Thermodynamics* consists of five lessons that cover the energy involved in bond breaking and bond formation, the heat of reaction, the heat of formation, Hess's law, entropy, and Gibb's free energy.

- **Lesson 27.1 Energy Change in Reactions**
- **Lesson 27.2 Enthalpy**
- **Lesson 27.3 Spontaneous Processes**
- **Lesson 27.4 Entropy**
- **Lesson 27.5 Gibb's Free Energy**

Overview

In these lessons, students will explore:

- Energy changes in endothermic and exothermic reactions.
- Enthalpy and Hess's Law of Heat Summation.
- Spontaneous and non-spontaneous events and reactions.
- The disorder of chemical systems.
- Gibb's Free energy and spontaneity.

Science Background Information

Pacing the Lessons

Use the table below as a guide for the time required to teach the lessons of *Thermodynamics*.

TABLE 27.1: **60 Minute Class Periods per Lesson**

Lesson	Number of Class Periods
27.1 *Energy Change in Reactions*	1.0
27.2 *Enthalpy*	2.0
27.3 *Spontaneous Processes*	1.0
27.4 *Entropy*	1.0
27.5 *Gibb's Free Energy*	1.5

Managing Materials

The following items are needed to teach the strategies and activities described in the Teachers Edition of the FlexBook for *Thermodynamics*.

TABLE 27.2: **Thermodynamics Materials List**

Lesson	Strategy or Activity	Materials Needed
27.1		
27.2		
27.3		
27.4		
27.5		

Multimedia Resources

You may find these additional internet resources helpful when teaching *Thermodynamics.*

Possible Misconceptions

Making the FlexBook Flexible

An important advantage of the FlexBook is the ability it gives you, the teacher, to select the chapters and lessons that you think are most important for your own classes. You should also consult the standards correlation table that follows when selecting chapters and lessons to include in the FlexBook for your classes.

TABLE 27.3: **Standard Addressed by the Lessons in Thermodynamics**

Lesson	California Standards	NSES Standards	AAAS Benchmarks
27.1			
27.2	7e		
27.3	7b		
27.4			
27.5	7f		

27.1 Energy Change in Reactions

Key Concepts

In this lesson students explore energy changes in endothermic and exothermic reactions.

Lesson Objectives

- Define energy, potential energy, kinetic energy.
- Define endothermic and exothermic reactions.
- Describe how heat is transferred in endothermic and exothermic reactions.

Lesson Vocabulary

energy The ability to do work.

potential energy The energy of position or stored energy.

kinetic energy The energy of motion.

enthalpy The amount of energy a system or substance contains.

heat The energy that is transferred between the system (reactants and products) and the surroundings.

temperature The average kinetic energy of the molecules of a substance.

Strategies to Engage

- Have students read the lesson objectives. Facilitate a discussion with students about what they already know about the key concepts to be explored in this lesson. Use this opportunity to gauge student understanding and address misconceptions.

Strategies to Explore

Strategies to Extend and Evaluate

- Have students work in groups to come up with a way to describe and explain endothermic and exothermic processes to elementary school students.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Review Questions that are listed at the end of each lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Define endothermic and exothermic reactions and give an example of each.
2. How does a campfire involve energy?
3. If a chemical reaction absorbs heat from the surroundings, it is said to be what?
   a. in equilibrium
   b. in a closed system
   c. an exothermic reaction
   d. an endothermic reaction
4. If a chemical reaction releases heat to the surroundings, it is said to be what?
   a. in equilibrium
   b. in a closed system
   c. an exothermic reaction
   d. an endothermic reaction
5. Symbolically, change in enthalpy is represented as:
   a. H
   b. $\triangle H$
   c. E
   d. $\triangle E$
6. Which of the following processes would be endothermic?
   a. natural gas burning
   b. melting chocolate
   c. fireworks exploding
   d. Steam condensing
7. Which of the following processes would be exothermic?
   a. gasoline burning

b. evaporation of ether
c. melting butter
d. boiling water

27.2 Enthalpy

Key Concepts

In this lesson students explore enthalpy and Hess's Law of Heat Summation.

Lesson Objectives

- Define and understand enthalpy of reaction.
- Calculate the enthalpy of reaction using $\Delta H_{rxn} = \Delta H_{products} - \Delta H_{reactants}$.
- Describe, interpret, and draw potential energy diagrams.
- Define and understand ΔH_f.
- Define Hess's Law.
- Calculate ΔH.

Lesson Vocabulary

activation energy The minimum amount of energy necessary for a reaction to take place.

potential energy diagrams Show endothermic chemical reaction; the activation of energy and the potential energy of the reactants.

enthalpy of formation The heat required to form one mole of a substance from its elements at standard temperature and pressure.

Hess's Law If multiple reactions are combined, the enthalpy $(\triangle H)$ of the combined reaction is equal to the sum of all the individual enthalpies.

Strategies to Engage

Strategies to Explore

- Ask students to look at Figure 27.3 and write a paragraph to describe the illustration in their own words.

Strategies to Extend and Evaluate

- Have students complete the lab ***Heat of Reaction-Hess's Law***. This lab is located in the Supplemental Lab Book.

Lesson Worksheets

Copy and distribute the lesson worksheets in the *CK-12 Chemistry Workbook* titled **Enthalpy Worksheet** and **Hess's Law Worksheet**. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 27.2 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Define the Hess's Law and the need for using this method.
2. Draw a potential energy diagram to represent the reaction below.

$$S_8(s) + 8\ Cl_2(g) \rightarrow 8\ SCl_2(s) \quad \triangle H = -376 \text{ kJ}$$

3. Which of the following does not have a $\triangle H_f = 0$?
   a. $H_2O_{(l)}$
   b. $O_{2(g)}$
   c. $H_{2(g)}$
   d. $Fe_{(s)}$
4. Which statement would describe an endothermic reaction?
   a. The potential energy of the reactants is greater than the potential energy of the products.
   b. The potential energy of the reactants is less than the potential energy of the products.
   c. Energy is released in the chemical reaction.
   d. The energy required to break bonds is more than the energy produced when bonds are formed.
5. Given the reaction: $2\ HCl(g) \rightarrow H_2(g) + Cl_2(g)$, $\triangle H = 185$ kJ, what would be the $\triangle H$ for the following reaction:$\frac{1}{2}H_2(g) + \frac{1}{2}Cl_2(g) \ \rightarrow \ HCl(g)$?
   a. 185
   b. -185
   c. 92.5
   d. -92.5
6. Which of the following reactions represents that for a $\triangle H_f$?
   a. $4\ Fe(s) + 3\ O_2(g) \rightarrow 2\ Fe_2O_3(s)$
   b. $SO_2(g) + \frac{1}{2}O_2(g) \rightarrow SO_3(g)$
   c. $2\ Al(s) + \frac{3}{2}O_2(g) \rightarrow Al_2O_3(s)$
   d. $\frac{1}{2}C_4H_{10}(g) + \frac{1}{2}H_2(g) \rightarrow C_2H_6(g)$

7. Hydrogen sulfide can mix with carbon dioxide to make a very smelly liquid, carbon disulfide. Given that the enthalpies of formation for $CO_2(g)$, $H_2S(g)$, $CS_2(l)$, and $H_2O(l)$ are -393.5 kJ/mol, -20.6 kJ/mol, 116.7 kJ/mol, and -285.8 kJ/mol, respectively, calculate $\triangle H_{\text{rxn}}$.
8. Ethene is a common compound used in the production of plastics for plastic bottles. Using the following data, calculate the $\triangle H_{\text{rxn}}$ for ethene. $2\ C(g) + 2\ H_2(g) \rightarrow C_2H_4(g)$

27.3 Spontaneous Processes

Key Concepts

In this lesson students explore spontaneous and non-spontaneous events and reactions.

Lesson Objectives

- Define a spontaneous and non-spontaneous reaction.
- Identify processes as either spontaneous or non-spontaneous.
- Describe how endothermic and exothermic reactions can be spontaneous or non-spontaneous.
- Explain the lack of correlation between spontaneity and speed of reaction.

Lesson Vocabulary

spontaneous event (or reaction) A change that occurs without outside inference; does not relate to rate of a reaction.

non-spontaneous event (or reaction) A change that will only occur with outside inference.

ionization A special type of dissociation reaction where a molecule ionizes in water to produce H^+ cations and the anion. Ionization reactions are specific to acids.

Strategies to Engage

Strategies to Explore

Strategies to Extend and Evaluate

- Have students write a paragraph comparing and contrasting the scientific and everyday definitions of spontaneity.
- In order to reinforce the fact that spontaneity does not relate to the speed of a reaction, have students research examples of spontaneous reactions that proceed slowly at room temperature. Students should be prepared to share their findings with the rest of the class.

Lesson Worksheets

There are no worksheets for this lesson.

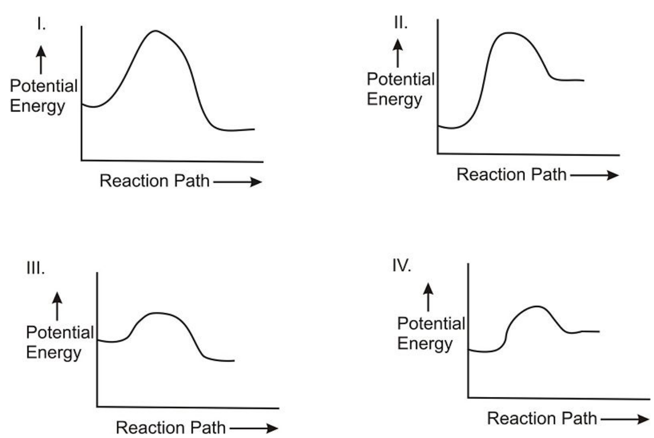
Review Questions

Have students answer the Lesson 27.3 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Distinguish between spontaneous and non-spontaneous reactions.
2. Why are spontaneous reactions usually exothermic (but still can be endothermic)?
3. Which of the following processes would be spontaneous?
   a. dissolving table salt
   b. climbing Mt. Everest
   c. separating helium from nitrogen in a mixture of gases
   d. none of these are spontaneous
4. Which of the following processes would be non-spontaneous?
   a. iron rusting in air
   b. ice melting at $10°C$
   c. a wild fire
   d. the reaction of CO_2 and H_2O
5. Which of the following reactions are spontaneous?



   a. I and II
   b. I and III
   c. II and IV
   d. Not enough information is given
6. If a reaction is spontaneous and fast, draw a likely potential energy diagram.

27.4 Entropy

Key Concepts

In this lesson students explore the disorder of chemical systems.

Lesson Objectives

- Define entropy.
- Calculate change in entropy from standard entropies of formation.
- Relate entropy to the tendency toward spontaneity.
- Describe the factors that affect the increase or decrease in disorder.

Lesson Vocabulary

entropy A measure of the disorder of a system.

Strategies to Engage

Strategies to Explore

- Have less proficient readers make a main ideas/details chart as they read the lesson. Instruct them to divide a sheet of paper down the middle and record the main ideas on the left side and the details for each main idea on the right side. Have students save their chart for reviewing lesson content. DI **Less Proficient Readers**

Strategies to Extend and Evaluate

- Have pairs of students create a lesson, using dominoes or cards as visual aids, to explain entropy in a way that young children can understand.
- Have students write a paragraph describing the relationship between entropy and changes of state, temperature, and the number of product/reactant particles.

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Entropy Worksheet**. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 27.4 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Define Entropy.
2. Give an everyday example of entropy.
3. Which of the following examples will result in an increase in entropy?
   a. $H_2O_{(L)} \rightarrow H_2O_{(s)}$
   b. $(NH_4)_2SO_{4(s)} \rightarrow 2\,NH_4{}^+{}_{(aq)} + SO_4{}^{2-}{}_{(aq)}$
   c. $H_2O_{(g)} \rightarrow H_2O_{(L)}$
   d. $Ag^{1+}{}_{(aq)} + Cl^-{}_{(aq)} \rightarrow AgCl_{(s)}$
4. Which of the following would have the greatest entropy?
   a. $CO_2(s)$
   b. $H_2O_{(g)}$
   c. $CCl_{4(L)}$
   d. $CHCl_{3(L)}$
5. From the following equations, select those that tend to be spontaneous.
   a. i, ii, and v
   b. i and iii
   c. iii and iv
   d. i, iv, and v
      a. $N_{2(g)} + O_{2(g)} \rightarrow N_2O_{5(g)} +$ heat
      b. $H_2O_{(s)} +$ heat $\rightarrow H_2O_{(L)}$
      c. $N_{2(g)} + O_{2(g)} \rightarrow 2\,NO_{(g)} \quad \triangle H = 180.6\text{ kJ/mol}$
      d. $C_6H_{12}O_{6(s)} + 6\,O_{2(g)} \rightarrow 6\,CO_{2(g)} + 6\,H_2O_{(L)} \quad \triangle H = -2802\text{ kJ/mol}$
      e. $CaCl_{2(s)} \rightarrow Ca^{2+}{}_{(aq)} + 2Cl^-{}_{(aq)} +$ heat
6. Calculate the entropy of the following reactions.
   a. $CH_3CH_2OH_{(L)} + 3\,O_{2(g)} \rightarrow 2\,CO_{2(g)} + 3\,H_2O_{(L)}$
   b. $2\,AsF_{3(L)} \rightarrow 2\,As_{(s)} + 3\,F_{2(g)}$

TABLE 27.4:

Compound	$\triangle S$	Compound	$\triangle S$
$CH_3CH_2OH(l)$	213. J/K × mol	$As(s)$	35.1 J/K × mol
$O_2(g)$	69.9 J/K × mol	$F_2(g)$	202.7 J/K × mol
$CO_2(g)$	160.7 J/K × mol	$AsF_3(L)$	181.2 J/K × mol
$H_2O(l)$	205.0 J/K × mol		

7. (3) Predict whether the entropy will be positive or negative for each of the following:

   a. $CO_{2(s)} \rightarrow CO_{2(g)}$
   b. $C_{12}H_{22}O_{11}(s) \rightarrow C_{12}H_{22}O_{11}(aq)$
   c. $2\,NO(g) + O_{2(g)} \rightarrow 2\,NO_{2(g)}$

27.5 Gibb's Free Energy

Key Concepts

In this lesson students explore Gibb's Free energy and spontaneity.

Lesson Objectives

- Define Gibbs Free Energy.
- Calculate Gibbs Free Energy given the enthalpy and entropy.
- Use Gibbs Free Energy to predict spontaneity.

Lesson Vocabulary

Gibbs free energy The maximum energy available to do useful work.

Strategies to Engage

Strategies to Explore

- Have students convert the information in Table 27.4 into a paragraph that contains the information.
- As a class create a flowchart that can be used to predict spontaneity based on the Gibbs free energy equation.
- Use the mathematical calculations in this lesson to reduce the reliance on language skills. As you go through each example problem, use them to explain the concepts explored in this lesson. DI **English Language Learners**
- Divide students into groups of three or four to work on problems in this lesson.

Strategies to Extend and Evaluate

- Have students create a concept map relating the terms and objectives of the concepts explored in this chapter.

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Enthalpy, Entropy, and Free Energy**. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 27.5 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Define Gibbs free energy.
2. Summarize the conditions of spontaneity according to Gibbs free energy equation.
3. For the reaction $C_2H_5OH(L) \rightarrow C_2H_5OH(g)$, $\triangle S = 122.0$ J/K $\times$ mol , and $\triangle H = 42.59$ kJ/mol at $25°C$. Which of the following statements is true?
   a. the reaction will always be spontaneous
   b. the reaction will always be non-spontaneous
   c. the reaction will be spontaneous only at high temperatures
   d. the reaction will be spontaneous only at high temperatures
4. For the reaction $C_6H_6(l) + 3\ H_2(g) \rightarrow C_6H_{12}(l)$, $\triangle S = -101.6$ J/K $\times$ mol , and $\triangle H = -205.4$ kJ/mol at $25°C$. Which of the following statements is true?
   a. the reaction will always be spontaneous
   b. the reaction will always be non-spontaneous
   c. the reaction will be spontaneous only at high temperatures
   d. the reaction will be spontaneous only at low temperatures
5. For the reaction $COCl_2(g) \rightarrow CO(g) + Cl_2(g)$, $\triangle H = 109.6$ kJ/mol and $\triangle S = 137.1$ J/K $\times$ mol. What is the value of $\triangle G$ at $25.0°C$?
   a. 68.7 kJ/mol
   b. 106 kJ/mol
   c. -3.32×10^3 kJ/mol
   d. -4.08×10^4 kJ/mol
6. Hydrazine, $N_2H_{4(L)}$, has an important use in the space industry as rocket fuel. The preparation of hydrazine is shown in the equation below. If the value of $\triangle S$ is -393.8 J/K$\cdot$ mol at $15°C$, what is the value of $\triangle G$?

$$N_2O(g) + 3\ H_2(g) \rightarrow N_2H_4(l) + H_2O(l) \qquad \triangle H = -317.0 \text{ kJ}$$

   a. 1.132×10^5 kJ/mol
   b. 5.590×10^3 kJ/mol
   c. -311.1 kJ/mol
   d. -203.5 kJ/mol
7. Which of the following regarding reaction spontaneity is true?
   a. A reaction with a positive $\triangle S^o$ will always be spontaneous.
   b. A reaction with a negative $\triangle H^o$ will always be spontaneous.
   c. A reaction with a positive $\triangle S^o$ and a negative $\triangle H^o$ will always be spontaneous.

d. A reaction with a negative $\triangle S^o$ and a negative $\triangle H^o$ will always be spontaneous.
e. A reaction with a positive $\triangle S^o$ and a positive $\triangle H^o$ will always be spontaneous.

8. Using the data provided in the table, find the values of (a) $\triangle H$, (b) $\triangle S$, and finally (c) $\triangle G$ at $100°C$. Is the system spontaneous or non-spontaneous?

$$CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(g)$$

TABLE 27.5:

	$\triangle H$	$\triangle S$
$CuO(s)$	-155.2 kJ/mol	43.5 J/K × mol
$H_2(g)$	0 kJ/mol	131.0 J/K × mol
$Cu(s)$	0 kJ/mol	33.3 J/K × mol
$H_2O(g)$	-241.8 kJ/mol	188.7 J/K × mol

Chapter 28 Electrochemical Cells TE

Chapter Outline

Unit 11 Electrochemistry

Outline

This unit, *Electrochemistry*, includes one chapter that covers oxidation-reduction and electrochemical cells.

- **Chapter 28 Electrochemistry**

Overview

Electrochemistry This chapter covers oxidation-reduction and electrochemical cells.

Electrochemistry

Outline

This chapter *Electrochemistry* consists of five lessons that cover oxidation-reduction and electrochemical cells.

- **Lesson 28.1 origin of the Term oxidation**
- **Lesson 28.2 oxidation-Reduction**
- **Lesson 28.3 Balancing Redox Equations Using the oxidation Number Method**
- **Lesson 28.4 Electrolysis**
- **Lesson 28.5 Galvanic Cells**

Overview

In these lessons, students will explore:

- The phlogiston and Lavoisier theories of combustion.
- oxidation, reduction, oxidation numbers, and oxidizing and reducing agents.
- The oxidation number method of balancing redox equations.
- Electrolysis and electrolysis apparatus.
- Redox reactions that occur in galvanic cells.

Science Background Information

This information is provided for teachers who are just beginning to instruct in this subject area.

Dependence of Cell Potential on Concentration

Non-Standard Cells

For the most part, the discussion of cell potential in high school chemistry deals with cells under standard conditions. Standard conditions for cells is $25^\circ\ C, 1.0\ atm$ pressure, and the concentrations of ions is $1.0\ M$. If you were building a cell to use to do work, you would not build a standard cell. Standard cells are used essentially for teaching or experimentation. They do not have either the maximum voltage, or the maximum capacity, that can be built into Galvanic cells. The advantage of standard cells is that their voltages are precisely predictable and easily calculated.

The cell potential for galvanic cells is closely related to the net movement of materials from reactants to products. A faster net reaction would produce a greater cell potential, and a slower net reaction would produce a smaller cell potential. At equilibrium, the net reaction is zero and therefore, the cell potential would be zero. If the forward reaction rate is increased with no change in the reverse reaction rate, then the net forward reaction is greater, and the cell potential will be greater. If the reverse reaction rate is decreased with no chain in the forward rate, then the net forward reaction is greater, and the cell potential will be greater.

Consider the cell composed of the standard half cells of aluminum and manganese.

$$2Al_{(s)} + 3Mn^{2+}_{(aq)} \rightarrow 2\,Al^{3+}_{(aq)} + 3Mn_{(s)} \qquad E^\circ_{cell} = 0.48\ V$$

The E°_{cell} for this reaction is determined when the concentrations of manganese ion and the aluminum ion are both $1.0\ M$. This is the voltage of this cell at standard conditions. If the concentration of the manganese ion is increases, the forward reaction rate will increase, and the net movement of material in the forward direction will increase. This increase in the net movement of material in the forward direction will cause the cell voltage to be higher. If the concentration of the aluminum ion is decreased, the reverse reaction rate will decrease, and the net movement of material in the forward direction will increase. This increase in the net movement of material in the forward direction causes the voltage to be higher. If the $[Mn^{2+}]$ concentration is decreased, the forward reaction rate will decrease, and the net movement of material in the forward direction will decrease. Therefore, the voltage of the cell will be lower. If the $[Al^{3+}]$ concentration is increased, the reverse reaction rate will be increased, and the net movement of material in the forward direction will decrease. Therefore, the voltage of the cell will be lower.

Cells that do not have the concentrations of ions at $1.0\ M$ are called non-standard cells. In the cell described above, $[Mn^{2+}] > 1.0\ M$ will cause the cell voltage to be greater than $0.48\ V$ and $[Al^{3+}] > 1.0\ M$ will cause the cell voltage to be less than $0.48\ volts$. Vice versa would be true if the concentrations were less than $1.0\ M$.

Cell voltages for non-standard cells can also be calculated using the Nernst Equation.

The Nernst Equation is $E = E^\circ - \left(\frac{0.0591}{n}\right)(logQ)$, where E is the voltage of the non-standard cell. E° would be the voltage of these reactants and products if they were a standard cell, n is the moles of electrons transferred in the balanced reaction, and Q is the reaction quotient. The reaction quotient is the equilibrium constant expression, but when the reaction is not at equilibrium it is called the reaction quotient.

The reaction quotient for the example cell used here is $Q = \frac{[Al^{3+}]^2}{[Mn^{2+}]^3}$.

The Nernst Equation for this cell is $E = E^\circ - \left(\frac{0.0591}{n}\right)\left(log\frac{[Al^{3+}]^2}{[Mn^{2+}]^3}\right)$.

If you follow the mathematics for the case when both ions concentrations are 1.0 M, the reaction quotient would be 1 and the log of 1 is zero. Therefore, the second term in the Nernst Equation is zero and $E = E^\circ$.

Let's take the case of the example cell when $[Mn^{2+}] = 6.0\ M$ and $[Al^{3+}] = 0.10\ M$.

$$E = E^\circ - \left(\frac{0.0591}{n}\right)\left(log\frac{[Al^{3+}]^2}{[Mn^{2+}]^3}\right)$$
$$E = 0.48\ V - \left(\frac{0.0591}{6}\right)\left(log\frac{[0.10]^2}{[6.0]^3}\right)$$
$$E = 0.48\ V - \left(\frac{0.0591}{6}\right)(-4.33)$$
$$E = 0.48\ V - (-0.04\ V)$$
$$E = 0.52\ V$$

Concentration Cells

If we attempt to construct a standard cell from the same two reactants, we do not get a reaction or a cell voltage. Suppose we attempt to build a cell with two silver half-cells.

$$Ag_{(s)} + Ag^+_{(aq)} \rightarrow Ag^+_{(aq)} + Ag_{(s)}$$

If this is a standard cell, the half-cell voltage for the oxidation half-reaction is $-0.80\ V$ and the half-cell voltage for the reduction half-reaction is $+0.80\ V$. Clearly the net voltage is 0. It is possible, however, to produce a voltage using the same two half-reactions if we alter the concentrations of the ions. Such a cell is called a concentration cell and its voltage can be calculated using the Nernst Equation.

The Nernst Equation would look like this:

$E = E^\circ - \left(\frac{0.0591}{n}\right)\left(log\frac{[Ag^+]}{[Ag^+]}\right)$, where the silver ion concentration in the numerator is the concentration of the silver ion in the products and the silver ion concentration in the denominator is the silver ion concentration in the reactants.

Suppose we build this cell using a concentration of silver ion in the products of 0.010 M and a silver ion concentration in the reactants of 6.0 M. The E° in this case is zero and $n = 1$.

$$E = E^\circ - \left(\frac{0.0591}{n}\right)\left(log\frac{[Ag^+]}{[Ag^+]}\right)$$
$$E = 0\ V - (0.0591)\left(log\frac{[0.010]}{[6.0]}\right)$$
$$E = 0\ V - (0.0591)(-2.78\ V)$$
$$E = 0\ V - (-0.16\ V)$$
$$E = 0.16\ V$$

Pacing the Lessons

Use the table below as a guide for the time required to teach the lessons of *Electrochemistry*.

TABLE 28.1: 60 Minute Class Periods per Lesson

Lesson	Number of Class Periods
28.1 *origin of the Term oxidation*	0.5
28.2 *oxidation-Reduction*	2.0
28.3 *Balancing Redox Equations Using the oxidation Number Method*	2.0
28.4 *Electrolysis*	2.0
28.5 *Galvanic Cells*	2.0

Managing Materials

The following items are needed to teach the strategies and activities described in the Teachers Edition of the FlexBook for *Electrochemistry*.

TABLE 28.2: Electrochemistry Materials List

Lesson	Strategy or Activity	Materials Needed
28.1		
28.2	Engagement Activity	Pre-1982 penny, 100 *mL* flask, concentration nitric acid
28.3		
28.4		
28.5		

Multimedia Resources

You may find these additional internet resources helpful when teaching *Electrochemistry*:

- Gummy Bear Terminator Demo: http://quiz2.chem.arizona.edu/preproom/Demo%20Files http://quiz2.chem.arizona.edu/preproom/Demo%20Files /gummi_bear_terminator.htm
- oxidation-reduction demos: http://sites.google.com/site/chemistrydemos/7–chemcial-reaction/oxidation-reduction
- Electrochenical Cell animation: http://chemmovies.unl.edu/ChemAnime/ECZCELLD/ECZCELLD.html
- Electroplating lesson: http://www.finishing.com/faqs/howworks.html

Possible Misconceptions

Making the FlexBook Flexible

An important advantage of the FlexBook is the ability it gives you, the teacher, to select the chapters and lessons that you think are most important for your own classes. You should also consult the standards correlation table that follows when selecting chapters and lessons to include in the FlexBook for your classes.

TABLE 28.3: Standard Addressed by the Lessons in Electrochemistry

Lesson	California Standards	NSES Standards	AAAS Benchmarks
28.1			
28.2	3g		
28.3	3g		
28.4			
28.5			

28.1 Origin of The Term Oxidation

Key Concepts

In this lesson students learn the phlogiston and Lavoisier theories of combustion.

Lesson objectives

- The students will define the term "oxidation."

Lesson Vocabulary

combustion A group of chemical reactants in which the reactants are fuel and oxygen gas.

phlogiston The "fire substance" from a former theory of combustion.

Strategies to Engage

- Review key concepts students have already explored that relate to this lesson by asking what they already know about combustion. Use this opportunity to gauge student understanding of the properties of liquids and to clear up any misconceptions.

Strategies to Explore

- Have less proficient readers make a main ideas/details chart as they read the lesson. Instruct them to divide a sheet of paper down the middle and record the main ideas on the left side, and the details for each main idea on the right side. Have students save their chart for reviewing lesson content. DI **Less Proficient Readers**

Strategies to Extend and Evaluate

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 28.1 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. What would be the experimental difference between phlogisticated air and de-phlogisticated air?

28.2 Oxidation-Reduction

Key Concepts

In this lesson students explore oxidation, reduction, oxidation numbers, and oxidizing and reducing agents.

Lesson objectives

- The students will assign the correct oxidation number to any element in a compound or ion.
- In an oxidation-reduction equation, the students will identify the substance being oxidized, the substance being reduced, the oxidizing agent, and the reducing agent

Lesson Vocabulary

oxidation

1. A chemical combination with oxygen (old definition).
2. A loss of electrons in an atom or an increase in the oxidation state of an atom (modern definition).

oxidation numbers In ionic compounds, it is equal to the ionic charge. In covalent compounds, it is the charge assigned to the atom in accordance with a set of rules.

oxidation state In ionic compounds, it is equal to the ionic charge. In covalent compounds, it is the charge assigned to the atom in accordance with a set of rules.

oxidizing agent A substance that gains electrons in a chemical reaction or undergoes an increase in its oxidation state.

reducing agent The substance in a redox reaction that loses electrons or increases its oxidation state.

reduction The gain of electrons or decrease in oxidation state in a chemical reaction.

Strategies to Engage

- Have students set up a two-column table. Have them label one column "oxidation" and the other column "reduction". As you explore the information in this lesson, have students write notes in the appropriate column.

- Perform a simple demonstration of a redox reaction. This demonstration must be performed in a fume hood. Add an old copper penny (pre-1982) to a 100 *mL* flask containing 30 *mL* of concentrated nitric acid. Nitric acid oxidizes copper metal to produce copper nitrate. Nitric acid is highly corrosive. The gas produced in this reaction is highly toxic.

Strategies to Explore

- It may be helpful for students to remember LEO the lion goes GER. LEO = loss of electrons is oxidation and GER = gain of electrons is reduction.
- Emphasize to students that although some reactions are referred as oxidation, reduction always accompanies oxidation.
- Have students write a paragraph to compare and contrast the old and new definitions of oxidation.

Strategies to Extend and Evaluate

- Encourage interested groups of students to create Keynote or PowerPoint slideshows explaining redox reactions, oxidizing and reducing agents, and assigning oxidation numbers to share with the rest of the class. Students should include and examples of each concept.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 28.2 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Indicate the oxidation numbers for each of the following atoms.
    a. Mn^{2+}
    b. Al
    c. Al in Al_2O_3
    d. Br in $NaBr$
    e. Fe in Fe_2O_3
    f. arsenic in AsO_4^{3-}
    g. chlorine in ClO_4^-
    h. sulfur in H_2SO_3
2. In the following reaction, identify the element that is being oxidized and the element that is being reduced. $MnO_2 + 4\ HCl \rightarrow MnCl_2 + Cl_2 + 2\ H_2O$

28.3 Balancing Redox Equations Using the Oxidation Number Method

Key Concepts

In this lesson students explore the oxidation number method of balancing redox equations.

Lesson objectives

- Given a redox reaction, the students will determine which substances are changing their oxidation state.
- Given a redox reaction, the students will balance the equation using the oxidation number method.

Lesson Vocabulary

Strategies to Engage

- Before beginning this lesson, review with students how to assign oxidation numbers and how to identify the substances in equation that are oxidized and reduced. Use this opportunity to gauge student understanding of the key concepts explored so far in this chapter, and to clear up any misconceptions.

Strategies to Explore

- Use the mathematical calculations in this lesson to reduce the reliance on language skills. As you go through each example problem, use them to explain the concepts explored so far in this lesson. DI **English Language Learners**

Strategies to Extend and Evaluate

- Have students complete the lab named ***An Activity Series Lab***. This lab is located in the Supplemental Lab Book.

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Balancing Redox Equations**. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 28.3 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Balance the following equation using the oxidation number method.

$$HNO_3 + Br_2 \rightarrow HBrO_3 + NO_2 + H_2O$$

2. Balance the following equation using the oxidation number method.

$$Zn + HNO_3 \rightarrow Zn(NO_3)_2 + NO_2 + H_2O$$

3. In terms of electron gain and loss, explain why chlorine undergoes both oxidation and reduction in the following unbalanced reaction.

$$Cl_2 + KOH \rightarrow KCl + KClO_3 + H_2O$$

4. Balance the equation in the previous problem.

28.4 Electrolysis

Key Concepts

In this lesson students explore electrolysis and electrolysis apparatus.

Lesson objectives

- Given a diagram of an electrolysis apparatus including the compound being electrolyzed, the students will identify the anode and the cathode.
- Given a diagram of an electrolysis apparatus including the compound being electrolyzed, the students will write the oxidation and reduction half-reactions.

Lesson Vocabulary

anode The electrode at which oxidation occurs.

battery A group of two or more cells that produces an electric current.

cathode The electrode at which reduction occurs.

electrolysis A chemical reaction brought about by an electric current.

electroplating A process in which electrolysis is used as a means of coating an object with a layer of metal.

Strategies to Engage

- Facilitate a discussion with students about their knowledge of electroplated objects such as jewelry. Use this opportunity to gauge student understanding, address misconceptions, and generate curiosity for the concepts explored in this lesson.

Strategies to Explore

- Use Figure 28.1 to explain as many concepts as possible. DI **English Language Learners**
- Perform the ***The Electrolysis of Water*** *and* **Copper to Silver to Gold** demonstrations. These demonstrations are located in the Supplemental Lab Book.

Strategies to Extend and Evaluate

- Have students do library research on the topic of real-life applications of electrochemistry such as antioxidant compounds, water purification and batteries. Then have students prepare a written report, Keynote or PowerPoint slideshow, or poster display.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 28.4 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Write the equations for the reactions that occur at the anode and at the cathode in the electrolysis of molten KBr.

28.5 Galvanic Cells

Key Concepts

In this lesson students explore redox reactions that occur in galvanic cells.

Lesson objectives

- The students will describe the conditions necessary for a cell to be standard cell.
- Given a table of standard reduction potentials and a diagram or description of a Galvanic cell, the students will balance the redox equation, calculate the standard cell potential, and determine the direction of electron flow in the external circuit.

Lesson Vocabulary

anode The electrode at which oxidation occurs.

cathode The electrode at which reduction occurs.

electrochemical cell An arrangement of electrodes and ionic solutions in which a spontaneous redox reaction is used to produce a flow of electrons in an external circuit.

salt bridge A U-shaped tube containing an electrolyte that connects two half-cells in an electrochemical cell.

voltage The potential difference between two points in an electric circuit.

Strategies to Engage

Strategies to Explore

- Ask students to look at Figure 28.3 and use write a paragraph to describe the illustration.

Strategies to Extend and Evaluate

- Have students play a review game called "Two Truths and a Lie" using what they know about electrochemistry. To do this, pair students, and have each pair write three statements, two of which are facts about electrochemistry, and one of which is a plausible "lie." Then have each pair join with two other pairs to share what they wrote and try to guess which of the statements are "lies" and which are "truths."
- Have students use grid paper to make a crossword puzzle using the vocabulary terms in this chapter. Ask students to exchange papers with a classmate and solve each other's puzzles.
- As a review of chapter vocabulary, have students divided a sheet of paper into three columns. Have students label the columns: Terms I know, Terms I think I know, and Terms I need to learn. Have them write the vocabulary terms in the appropriate column. Have them attempt to define the vocabulary terms in the first two columns, then check their answers. Instruct students to create flash cards for terms in the last columns along with any terms from the first two columns that they did not define correctly.

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Electrochemical Cells**. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Review Questions

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. In a standard cell composed of the zinc and copper half-cells, will the current in the external circuit flow from Zn to Cu or from Cu to Zn?
2. For a standard cell with the following balanced equation. $Sn_{(s)} + 2\,Ag^+ \rightarrow Sn^{2+} + 2\,Ag_{(s)}$
   a. what is being oxidized?
   b. what is the reducing agent?
   c. what is the $E^\circ{}_{\text{NET}}$?
   d. after the reaction has reached equilibrium, what will be its voltage?
3. Balance the following equation using the half-reactions from the standard reduction potential chart. $Cr_2O_7{}^{2-} + Fe^{2+} + H^+ \rightarrow Cr^{3+} + Fe^{3+} + H_2O$

CHAPTER 29 Radioactivity and the Nucleus TE

Chapter Outline

Unit 12 Nuclear Chemistry

Outline

This unit, Nuclear Chemistry, includes one chapter that is an introduction to radioactivity, nuclear equations, and nuclear energy.

- **Chapter 29 Nuclear Chemistry**

Overview

Nuclear Chemistry This chapter is an introduction to radioactivity, nuclear equations, and nuclear energy.

Nuclear Chemistry

Outline

The chapter *Nuclear Chemistry* consists of seven lessons that serve as an introduction to radioactivity, nuclear equations, and nuclear energy.

- **Lesson 29.1 The Discovery of Radioactivity**
- **Lesson 29.2 Nuclear Notation**

- **Lesson 29.3 Nuclear Force**
- **Lesson 29.4 Nuclear Disintegration**
- **Lesson 29.5 Nuclear Equations**
- **Lesson 29.6 Radiation Around Us**
- **Lesson 29.7 Applications of Nuclear Energy**

Overview

In these lessons, students will explore:

- The discovery of radioactivity and common emissions from naturally radioactive nuclei.
- Nuclear symbols and the information contained in them.
- The relationship between nuclear force and nuclear energy.
- Radioactive decay.
- Equations for nuclear transmutations.
- Common nuclear emissions and half-life.
- Uses of radiation and nuclear energy.

Science Background Information

This information is provided for teachers who are just beginning to instruct in this subject area.

Carbon-14 Dating

If you're a fan of the television CSI shows or other mystery or crime programming, you are probably aware of different means to estimate the timing of the poor unfortunate's demise, dealing with factors such as body temperature, etc. For an archaeologist or an anthropologist, the trail of evidence is much colder. In the late 1940's, Willard Libby of the University of Chicago devised a method to establish the age of even the oldest unearthed fossils based on the remaining amount of radioactive ^{14}C. This isotope is one of three for the element carbon, which is ubiquitous in living systems. Carbon has an atomic number of six, and exists in three nuclear configurations: with six, seven and eight neutrons respectively. Thus the ^{14}C isotope has a nucleus consisting of six protons and eight neutrons. This assembly spontaneously decays into Nitrogen-14 and the release of beta radiation.Radioisotopic carbon has been measured to decay at a constant rate, with half the initial amount remaining, after 5730 years. If it is assumed that the ^{14}C is not replaced, the loss of ^{14}C suggests the time interval since the artifact last exchanged CO_2 with the atmosphere.

When the ratio of the remaining amount of ^{14}C to ^{12}C is compared to the same ratio in a living organism, the amount of time elapsing since the organism's death can be analyzed.

Thus, over time, in any material that contains the element carbon, the amount of remaining ^{14}C in a sample is an indicator of the age of the artifact.

One of the best-known applications of this technique was in the analysis of Ötzi, the alpine Ice Man. Found in a region straddling the Austrian-Italian border, by hikers in 1991, Ötzi was the name given to the partially mummified remains of a hunter located still frozen into a glacier. Ötzi provided a wealth of information to anthropologists in that he was still dressed with fur boots, and a pack with tools including a copper hatchet and arrows. Researchers even discovered the menu of his last meal, probably deer meat and wheat bran, by analyzing his stomach contents. Analysis of small tissue samples from his corpse suggest that he lived from 5300 to 5100 years ago, before the Bronze Age. The construction and contents of his clothing and possessions provide an invaluable insight into the culture and technological sophistication of his age.

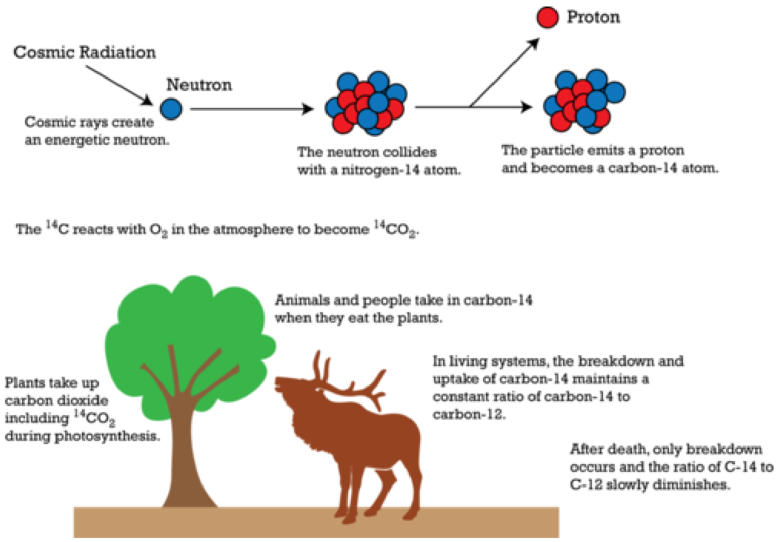


FIGURE 29.1

Other applications of this technology include dating of the Dead Sea scrolls, and analysis of the time period of the cave art found in central Europe. Ocean sediment samples, and even a meteorite believed to have originated on Mars!

Pacing the Lessons

Use the table below as a guide for the time required to teach the lessons of *Nuclear Chemistry*.

TABLE 29.1: 60 Minute Class Periods per Lesson

Lesson	Number of Class Periods
29.1 *The Discovery of Radioactivity*	1.0
29.2 *Nuclear Notation*	1.0
29.3 *Nuclear Force*	1.0
29.4 *Nuclear Disintegration*	1.5
29.5 *Nuclear Equations*	1.5
29.6 *Radiation Around Us*	1.0
29.7 *Applications of Nuclear Energy*	1.0

Managing Materials

The following items are needed to teach the strategies and activities described in the Teachers Edition of the FlexBook for *Nuclear Chemistry*.

TABLE 29.2: Nuclear Chemistry Materials List

Lesson	Strategy or Activity	Materials Needed
29.1	Exploration Activity	Self-developing film, various sources of low-level radiation.
29.2		
29.3		
29.4		
29.5	Exploration Activity	Dominoes

Possible Misconceptions

Identify: Students may think that radiation is harmful in all forms.

Clarify: Radiation is the transfer of radiant energy by means of electromagnetic waves. The word radiation is sometimes used to describe the energy that travels in the form of electromagnetic waves (radiant energy). Although some forms of radiant energy are harmful, most are not. Radiation has many uses such as energy generation, industry, and medicine.

Making the FlexBook Flexible

An important advantage of the FlexBook is the ability it gives you, the teacher, to select the chapters and lessons that you think are most important for your own classes. You should also consult the standards correlation table that follows when selecting chapters and lessons to include in the FlexBook for your classes.

TABLE 29.3: **Standard Addressed by the Lessons in Nuclear Chemistry**

Lesson	California Standards	NSES Standards	AAAS Benchmarks
29.1			
29.2	11c		
29.3	11a, 11b, 11g		
29.4	11d, 11e, 11g		
29.5	11b		
29.6	1i, 11c, 11e, 11f		
29.7	1f, 11b		

29.1 Discovery of Radioactivity

Key Concepts

In this lesson students explore the discovery of radioactivity and common emissions from naturally radioactive nuclei.

Lesson Objectives

- The students will describe the roles played by Henri Becquerel and Marie Curie played in the discovery of radioactivity.
- The students will list the most common emissions from naturally radioactive nuclei.

Lesson Vocabulary

alpha particle An alpha particle is a helium−4 nucleus.

beta particle A beta particle is a high speed electron, specifically an electron of nuclear origin.

gamma ray Gamma radiation is the highest energy on the spectrum of electromagnetic radiation.

Marie Curie Marie Curie was a physicist and chemist of Polish upbringing, and subsequently, French citizenship; a pioneer in the field of radioactivity, and the only person to ever win two Nobel prizes in science.

Strategies to Engage

- Review with students the structure of the atom by drawing a model on the board. Point out to students that chemical reactions involve electrons. Explain to students that in this chapter, they will explore nuclear reactions. Tell students that the focus will shift from the electrons to the nucleus, because nuclear reactions involve changes within the nucleus of the atom.
- Ask students what they remember about Rutherford's gold foil experiment. Explain to students that Rutherford knew that alpha particles were emitted by radioactive material, and that in this lesson they will learn more about these and other emissions from naturally radioactive nuclei.
- Students are likely to have heard about radiation in advertising and popular media (e.g., medicine, comic books, cartoons). Call on volunteers to share with the class anything they already know about radiation. Point out correct responses and clear up any misconceptions. Tell students they will learn more about radiation in this chapter.

Strategies to Explore

- Have students create a chart summarizing the properties of alpha particles, beta particles, and gamma rays explored in this lesson. Have students add information explored in the rest of the chapter.
- Have students write a newspaper article announcing the discovery of radioactivity. Encourage them to use factual information explored so far in this lesson.
- You can model the discovery of radioactivity by performing this simple demonstration. Cover self-developing film with heavy black paper to prevent visible light from exposing the film. Place an item that is a source of low-level radiation such as an ionizing smoke detector, antique ceramic dishware, a radium-dial watch face, or a weighted tape dispenser on top of the film, and put it in a location where it can be left undisturbed for several days.

Strategies to Extend and Evaluate

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 29.1 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Put the letter of the matching phrase on the line preceding the number. ___ 1. alpha particle - a. high energy electromagnetic radiation ___ 2. beta particle - b. a high speed electron ___ 3. gamma ray - c. a helium nucleus

29.2 Nuclear Notation

Key Concepts

In this lesson students explore nuclear symbols and the information contained in them.

Lesson Objectives

- The students will state the information contained in the atomic number of a nucleus.
- The students will state the information contained in the mass number of a nucleus.
- The students will subtract the atomic number from the mass number to determine the number of neutrons in a nucleus.
- Students will read and write complete nuclear symbols (know the structure of the symbols and understand the information contained in them).

Lesson Vocabulary

atomic number The atomic number indicates the number of protons in the nucleus.

mass number The mass number indicates the number of protons plus the number of neutrons in the nucleus.

electron An electron is a fundamental sub-atomic particle that carries a negative charge.

neutron A neutron is a sub-atomic particle with no electric charge and a mass slightly larger than a proton.

proton A proton is a fundamental sub-atomic particle with a net positive charge.

nucleus The nucleus of an atom if the very dense region, consisting of nucleons (proton and neutrons) at the center of an atom.

nuclei Nuclei is the plural of *nucleus*.

nucleon A nucleon is a constituent part (proton or neutron) of an atomic nucleus.

nuclide A type of nucleus specified by its atomic number and mass number.

Strategies to Engage

- Have students look at Figure 29.1. Review with students the information contained in a nuclear symbol, the definitions of both mass numbers and atomic numbers, and how to use the nuclear symbol to determine the number of protons, neutrons, and electrons in an atom of an element. Use this opportunity to gauge student understanding and address misconceptions about the concepts explored in this lesson.
- Preview the lesson vocabulary to find out what your students already know about the concepts to be explored in this lesson. Have students define each vocabulary term. At the end of the lesson encourage students to go back and write the correct definition for each incorrect definition.

Strategies to Explore

Strategies to Extend and Evaluate

- As a review of lesson content, have students create a ten-question quiz about the key concepts explored in this lesson. Have students exchange papers with another student who then takes the quiz. Have them hand the papers back to the original student who will assign a grade. Encourage students to discuss any incorrect answers.
- Have students write a one-paragraph summary of this lesson. Instruct students to correctly use each vocabulary term at least one time in the summary.

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Nuclear Notation**. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 29.2 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Write the complete nuclear symbol for a nucleus of chlorine that contains 17 protons and 20 neutrons.
2. Write the complete nuclear symbol for a nucleus of oxygen that contains 8 protons and 10 neutrons.
3. If a nucleus of uranium has a mass number of 238, how many neutrons does it contain?
4. In the nuclear symbol for a beta particle, what is the atomic number?
5. Is it possible for isotopes to be atoms of different elements? Explain why or why not.
6. How many neutrons are present in a nucleus whose atomic number is one and whose mass number is one?
7. Name the element of an isotope whose mass number is 206 and whose atomic number is 82.
8. How many protons and how many neutrons are present in a nucleus of lithium-7?
9. What is the physical difference between a $U-235$ atom and a $U-238$ atom?

10. What is the difference in the chemistry of a $U-235$ atom and a $U-238$ atom?

29.3 Nuclear Force

Key Concepts

In this lesson students explore nuclear force and nuclear energy.

Lesson Objectives

- Students will compare the energy released per gram of matter in nuclear reactions to that in chemical reactions.
- Students will express the equation for calculating the change in mass during nuclear reactions that is converted into energy.
- Students will express the relationship between nuclear stability and the nuclei's binding energy per nucleon ratio.

Lesson Vocabulary

binding energy Binding energy is the amount of energy that holds a nucleus together, and therefore, also the amount of energy required to decompose a nucleus into its component nucleons.

mass defect Mass defect is the difference between the sum of the masses of the nuclear components and the mass of the corresponding nucleus. Much of this lost mass is converted into binding energy.

nucleon Nucleon is a collective name for neutrons and protons.

Strategies to Engage

- Ask students if they have ever wondered how the protons and neutrons are able to remain together in the nucleus when neutrons are neutral, protons are positive, and like charges repel. Explain to students that in this chapter they will learn how this is possible. You may want to allow students to come up with their own explanations and discuss their ideas as a class.

Strategies to Explore

- On the board or chart paper, outline the main concepts of the lesson as a class. Discuss the main concepts as you prepare the outline.

Strategies to Extend and Evaluate

- Have students do library research on Einstein's equation, $E = mc^2$ and prepare a written report, Keynote or PowerPoint slideshow, or poster display.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 29.3 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Iron−56 is a very stable nucleus while cobalt−60 is an unstable nucleus. Which nucleus would you expect to have more binding energy per nucleon?
2. Calculate the mass defect and binding energy for a mole of carbon−14 given the data below.

The molar mass of carbon−14 is 14.003241 g/mol.

The molar mass of a proton is 1.007825 g/mol.

The molar mass of a neutron is 1.008665 g/mol.

Mass in kilograms is converted into energy in Joules by multiplying the mass times the speed of light squared, $E = mc^2$.

The speed of light is 3.00×10^8 m/s.

$\text{kg} \cdot \text{m}^2/\text{s}^2 = \text{Joules}$

29.4 Nuclear Disintegration

Key Concepts

In this lesson students explore radioactive decay.

Lesson Objectives

- Students will list some naturally occurring isotopes of elements that are radioactive.
- Students will describe the three most common emissions during natural nuclear decay.
- Students will express the changes in the atomic number and mass number of radioactive nuclei when an alpha particle is emitted.
- Students will express the changes in the atomic number and mass number of radioactive nuclei when a beta particle is emitted.
- Students will express the changes in the atomic number and mass number of radioactive nuclei when a gamma ray is emitted.
- Students will express that protons and neutrons are not indivisible and are composed of particles called quarks.
- Students will express the number of quarks that make up a proton or neutron.

Lesson Vocabulary

alpha decay Alpha decay is a common mode of radioactive decay in which a nucleus emits an alpha particle (a helium-4 nucleus).

beta decay Beta decay is a common mode of radioactive decay in which a nucleus emits beta particles. The daughter nucleus will have a higher atomic number than the original nucleus.

quark Quarks are physical particles that form one of the two basic constituents of matter. Various species of quarks combine in specific ways to form protons and neutrons, in each case taking exactly three quarks to make the composite particle.

Strategies to Engage

- Have students read the lesson objectives. Ask students to write down and try to complete each objective. Instruct students to use a scale of 1-5 to record how sure they are that they have correctly completed each objective (1= not, 5 = very sure). As you explore this lesson, encourage students to change their answers as necessary.

Strategies to Explore

- Have less proficient readers make a main ideas/details chart as they read the lesson,. Instruct them to divide a sheet of paper down the middle and record the main ideas on the left side and the details for each main idea on the right side. Have students save their chart for reviewing lesson content. DI **Less Proficient Readers**

Strategies to Extend and Evaluate

- Have a group of interested students do library research on how quarks were named. Students should be prepared to share their findings with the rest of the class.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 29.4 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Write the nuclear equation for the alpha decay of radon-198.
2. Write the nuclear equation for the beta decay of uranium-237.
3. There are six known quarks. The experimenter who discovers particles in nuclear physics gets the right to name the new particle. This has resulted in some very fanciful names for quarks. The six quarks are named, **up** quarks, **down** quarks, **charmed** quarks, **strange** quarks, **top** quarks, and **bottom** quarks. (The top and bottom quarks were originally named *truth* and *beauty* quarks, but the names were changed for some reason.) Protons and neutrons are each made of only up and down quarks and they are made of three quarks each. The up quark carries a charge of $+\frac{2}{3}$ and the down quark carries a charge of $-\frac{1}{3}$. Determine by the final charge on the proton and neutron, what combination of three up and down quarks are required to make a proton and what combination will make a neutron?

29.5 Nuclear Equations

Key Concepts

In this lesson students explore equations for nuclear transmutations.

Lesson Objectives

- The students will give definitions and examples of fission and fusion.
- The students will classify nuclear reactions as fission or fusion.
- Given a nuclear equation with one species missing, the student will be able to correctly fill in the missing particle.
- Students will write balanced equations for nuclear transmutations.

Lesson Vocabulary

artificial radioactivity Induced radioactivity that is produced by bombarding an element with high-velocity particles.

chain reaction A multi-stage nuclear reaction that sustains itself in a series of fissions in which the release of neutrons from the splitting of one atom leads to the splitting of others.

critical mass The smallest mass of a fissionable material that will sustain a nuclear chain reaction at a constant level.

fission A nuclear reaction in which a heavy nucleus splits into two or more smaller fragments, releasing large amounts of energy.

fusion A nuclear reaction in which nuclei combine to form more massive nuclei with the simultaneous release of energy.

natural radioactivity The radioactivity that occurs naturally, as opposed to induced radioactivity. Also known as spontaneous fission.

Strategies to Engage

- Have students read the review questions at the end of this section. This way, students will be familiar with the types of information that they will explore in this section.

Strategies to Explore

- Point out to students that although mass and atoms are not conserved in nuclear reactions like they are in chemical reactions, total mass number and total atomic number are conserved. So, in a nuclear equation the sum of the mass numbers of the reactants must equal the sum of the mass numbers of the products, and the sum of the charges of the reactants must equal the sum of the charges of the products.
- Have students write a short lesson comparing and contrast nuclear and chemical equations. Instruct students to include examples of each type of equation.
- Have students create a Venn diagram comparing and contrasting nuclear fusion and fission.
- Give groups of students a set of dominoes and challenge them to set up the dominoes in a way that will most closely model a nuclear chain reaction. The correct arrangement would resemble a triangle, where knocking down one domino would cause two to fall, and those two would each cause more to fall and so on.

Strategies to Extend and Evaluate

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Nuclear Equations**. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 29.5 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Only one particle is missing from this equation. What are its atomic and mass numbers?

$$^{14}_{7}\mathbf{N} +$$

*beginmatrix***4**
2

He $\rightarrow$
*beginmatrix***1**
H + ?

To what element does the missing particle in question #1 belong?

When a $U-235$ nucleus is struck by a neutron, the nucleus may be split into $Ce-144$ and $Sr-90$ nuclei, also emitting four electrons and two neutrons. Write the equation for this nuclear reaction.

Complete the following nuclear equation by supplying the missing particle.

$$^{196}_{85}\mathbf{At} \rightarrow$$

$beginmatrix\mathbf{4}$

$\mathbf{2}$

He + ?

Complete the following nuclear equation by supplying the missing particle.

$$^{208}_{84}\mathbf{Po} \rightarrow$$

$beginmatrix\mathbf{4}$

$\mathbf{2}$

He + ?

Complete the following nuclear equation by supplying the missing particle.

$$^{210}_{86}\mathbf{Rn} \rightarrow$$

$beginmatrix\mathbf{4}$

$\mathbf{2}$

He + ?

Complete the following nuclear equation by supplying the missing particle.

$$^{201}_{80}\mathbf{Hg} + ? \rightarrow$$

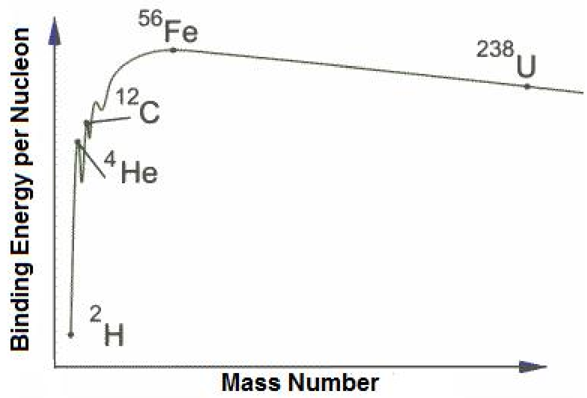
$beginmatrix\mathbf{201}$

$\mathbf{79}$

Au

Complete the following nuclear equation by supplying the missing particle.

$$? + \rightarrow {}^{210}_{84}\mathbf{Po} + {}^{0}_{-1}\mathbf{e}$$



Use information in the chart above to decide if carbon−12 nuclei were to be transmuted into other nuclei that were more stable, would this more likely be accomplished by fission or by fusion?

29.6 Radiation Around Us

Key Concepts

In this lesson students explore common nuclear emissions and half-life.

Lesson Objectives

- Students will calculate the amount of radioactive material that will remain after an integral number of half-lives.
- Students will describe how carbon-14 is used to determine the age of carbon containing objects.
- Students will qualitatively compare the ionizing power and penetration power of α, β, and γ particles.

Lesson Vocabulary

background radiation Radiation that comes from environment sources including the earth's crust, the atmosphere, cosmic rays, and radioisotopes. These natural sources of radiation account for the largest amount of radiation received by most people.

half-life The half-life of a radioactive substance is the time interval required for a quantity of material to decay to half its original value.

Strategies to Engage

- Students are likely to have heard about radiation in advertising and popular media (e.g., medicine, comic books, cartoons). Call on volunteers to share with the class anything they already know about radiation. Point out correct responses and clear up any misconceptions. Tell students they will learn more about radiation in this chapter.

Strategies to Explore

- Use pennies to model half-life. Place 50 pennies in a bag or cup. Shake for 10 *seconds*. Gently pour out the pennies. Count the number of pennies that are heads up. Explain to students that these are the "decayed" atoms. Return only the pennies that are tails up to the bag and shake for 10 *seconds*. Count the number of pennies that are heads up. Explain to students that these are the "decayed" atoms. Continue, shaking and counting until all the atoms have decayed. Instruct students to prepare a graph of number of decayed

atoms vs. time. Have students write a paragraph explaining how this activity relates to isotopes, half-life, and radioactivity.

Strategies to Extend and Evaluate

- Before the health effects of radiation were known and understood, products that contain radioactive isotopes such as radium were thought to be good for you. Have students research some of these products and present their findings to the class.

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Radioactivity**. Ask students to complete the worksheets alone or in pairs as a review of lesson content.

Review Questions

Have students answer the Lesson 29.6 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Which of the three common emissions from radioactive sources requires the heaviest shielding?
2. The half-life of radium−226 is about 1600 years. How many grams of a 2.00 gram sample will remain after 4800 years?
3. Sodium−24 has a half-life of about 15 hours. How much of an 16.0 grams sample of sodium-24 will remain after 60.0 hours?
4. A radioactive isotope decayed from 24.0 grams to 0.75 grams in 40.0 years. What is the half-life of the isotope?
5. What nuclide is commonly used in the dating of organic artifacts?
6. Why does an ancient wood artifact contain less carbon−14 than a piece of lumber sold today?
7. The half-life of $C-14$ is about $5,700$ years. An organic relic is found to contain $C-14$ and $C-12$ in a ratio that is about one-eighth as great as the ratio in the atmosphere. What is the approximate age of the relic?
8. Even though gamma rays are much more penetrating than alpha particles, it is the alpha particles that are more likely to cause damage to an organism. Explain why this is true.
9. The radioactive isotope calcium−47 has been used in the study of bone metabolism; radioactive iron−59 has been used in the study of red blood cell function; iodine−131 has been used in both diagnosis and treatment of thyroid problems. Suggest a reason why these particular elements were chosen for use with the particular body function.

29.7 Applications of Nuclear Energy

Key Concepts

In this lesson students explore uses of radiation and nuclear energy.

Lesson Objectives

- The students will trace the energy transfers that occur in a nuclear power reactor from the binding energy of the nuclei to the electricity that leaves the plant.
- The students will define the term "breeder reactor."
- The students will list some medical uses of nuclear energy.

Lesson Vocabulary

control rods Control rods are made of chemical elements capable of absorbing many neutrons and are used to control the rate of a fission chain reaction in a nuclear reactor.

cyclotron A cyclotron is a type of particle accelerator.

fall out Fall out is radioactive dust hazard from a nuclear explosion, so named because it "falls out" of the atmosphere where it was spread by the explosion.

fissile A fissile substance is a substance capable of sustaining a chain reaction of nuclear fission.

fissionable A fissionable material is material capable of undergoing fission.

Geiger counter A Geiger counter is an instrument used to detect radiation, usually alpha and beta radiation, but some models can also detect gamma radiation.

isotope Nuclei with the same number of protons but different numbers of neutrons.

linear accelerator A linear accelerator is a linear electrical device for the acceleration of subatomic particles.

moderator A neutron moderator is a medium which reduces the velocity of fast neutrons; commonly used moderators are regular (light) water, solid graphite, and heavy water.

nuclear pile A nuclear pile is a nuclear reactor.

Strategies to Engage

Strategies to Explore

- Have teams of students debate the use of nuclear reactions as an alternative source of energy.

Strategies to Extend and Evaluate

- Have students research the Chernobyl disaster's effect on biological systems. Students should prepare a written report of their findings.
- Students are likely to have heard about radioactivity in popular media. Have students bring in examples of bad science related to radioactivity from the web or from books, including comic books. Have them quote the claim, reference the source, and explain what is wrong.
- Have students read the ***Facts and Myths About Nuclear Power Plants*** extra reading located in the Supplemental Workbook.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 29.7 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. What is the primary physical difference between a nuclear electricity generating plant and a coal-burning electricity generating plant?
2. What do the control rods in a nuclear reactor do and how do they do it?
3. What is a breeder reactor?
4. Name two types of particle accelerators.
5. In the medical use of radioactivity, what does EBT stand for?
6. Is it possible for a nuclear explosion to occur in a nuclear reactor? Why or why not?

Chapter 29 Enrichment

Extra Readings

Facts and Myths About Civilian Nuclear Power Plants

- MYTH: Nuclear reactors may undergo a nuclear explosion killing tens of thousands of people.
- FACT: Civilian nuclear power plants in the U.S. never contain a supercritical mass of fissionable material and therefore, cannot explode even if operators tried to make them explode.
- MYTH: Nuclear power plants are not safe.
- FACT: 1. The radiation levels measured outside of the containment building of nuclear power plants are essentially the same as background radiation. 2. There was a nuclear accident in 1986 at the Chernobyl nuclear installation in the Soviet Union that resulted in the immediate death of 28 people (mostly employees and fire fighters), the subsequent death of 19 people, and 9 deaths from thyroid cancer apparently due to the accident. The number of injuries due to fall-out radiation from the accident in unknown (at that time, news from the Soviet Union was highly censored). Predictions of numbers of injuries by UNSCEAR (United Nations Scientific Committee on the Effects of Atomic Radiation) were in excess of 4,000 but the number was disputed by the IAEA (International Atomic Energy Agency).[5,6]

In either a nuclear explosion or a severe nuclear accident, a large amount of radioactive material rises into the upper atmosphere. At some later time, this radioactive material falls back to the earth, usually downwind from where the explosion or accident occurred. This material that falls back to the earth is known as **fall-out**. It is known that the fall-out from the Chernobyl accident was 30 times the fall-out from the Hiroshima and Nagasaki bombings[7] and that 300,000 people were evacuated from the Chernobyl area. It is also known that the fall-out reached as far as Sweden because Swedish workers at a Swedish nuclear power plant began testing positive for radiation, and after a thorough check of their own plant, officials determined the radiation was coming from fall-out from the Chernobyl plant. The greatest threat from fall-out is to children because children are growing and developing rapidly and radiation like all toxic materials have the greatest effect on children. It may still be years before all the effects of the Chernobyl accident are known. A small area called the Chernobyl Exclusion Zone is still closed, but the rest of the fall-out area is now considered safe.[7]

The Chernobyl accident occurred in an early model of Soviet reactor that had no reaction vessel and no containment building. The fuel rods and control rods were inserted into graphite blocks. The graphite blocks worked well as a moderator, but graphite is combustible. When a fire started in the reactor, the employees and emergency workers were unable to control it, and the graphite burned away releasing radiation into the environment. There are some military reactors of this type in the United States, such as the one at Hanford, WA and there have been some radiation injuries at that facility. Civilian nuclear power plants in the U.S., however, do not use the graphite block reactor design. All U.S. civilian nuclear generating plants use LWR reactors, which have the reactor core submerged in a vessel of water and surrounded by a containment building. They also have a series of fail-safe shut down safety measures.

- MYTH: Hundreds of uranium miners die every year from radiation sickness.
- FACT: Hundreds of uranium miners died in the early days of uranium mining but that problem was solved long ago. On the other hand, 5,000 coal miners continue to die every year worldwide due to cave-ins, explosions, and black lung disease.
- MYTH: Nuclear reactors produce a large amount of radioactive waste that will be dangerous for thousands of years.
- FACT: When the percentage of $U-235$ in fuel rods gets below a certain level, they will no longer function as fuel and must be replaced. Even though the radioactivity is too low to function as fuel rods, they are still

extremely dangerous, and must be isolated for a long period of time. Several suggestions have been made for storage of this used fuel, but even though the method is considered safe by nuclear scientists, the people who live in the area where the waste is to be stored strongly oppose having the material stored near them. At the present time, the used fuel rods are still submerged in the reaction vessels where they were replaced. Now that the US government has removed the ban on recycling used fuel, the amount of radioactive waste will be reduced to approximately one-fourth of the present amount. Not only will recycling help with the waste disposal problem, it will also reduce the cost of fuel. The Department of Energy is considering space disposal (rocket the waste into the sun), geological disposal (burying the waste thousands of feet underground in geologically stable areas), transmutation disposal (building a nuclear reactor that will consume nuclear waste; this idea was banned by President Carter, reinstated by President Reagan, and was being investigated by President Bush). For a complete discussion on the handling of nuclear waste, the internet has several dozen sites.

- MYTH: Nuclear reactors are particularly vulnerable to terrorist attack.
- FACT: In 1988, Sandia National Laboratories conducted a test by slamming a military F-4 Phantom jet fighter into a concrete block built to simulate a nuclear reactor containment building. The airplane hit the block at 481 *miles* per hour and while the airplane was demolished, the six-foot thick wall suffered a dent 2.5 *inches* deep[8]. The Turkey Point Nuclear Generating station (near Miami, Florida) suffered a direct hit by hurricane Andrew in 1992. Turkey Point has two fossil fuel units and two nuclear units. The fossil fuel plants suffered $90 million of damage while the nuclear containment buildings were undamaged.
- COMPLAINT: Some countries may divert nuclear reactor materials to weapon building.
- FACT: True.
- COMPLAINT: When the cooling water from nuclear power plants is dumped back into the original source (river, lake, bay), the temperature of the water over a period of time may be raised several degrees. The amount of oxygen that water will hold in dissolved form is highly dependent on the temperature. Active fish (so-called sport fish) frequently move away from areas where the temperature has increased a few degrees and less active fish (so-called trash fish) move in. To keep everyone happy, the use of cooling towers needs to be greatly increased so that the water returned to its natural source is at the same temperature as when it was taken.
- FACT: True.

29.8 References

1. Richard Parsons. CK12Foundation. CC-BY-SA

Chapter 30 The Chemistry of Carbon TE

Chapter Outline

The Chemistry of Carbon TE

Outline

This unit, *Organic Chemistry*, includes one chapter that introduces the structure and nomenclature of straight chain hydrocarbons, aromatic hydrocarbons, and the functional groups of hydrocarbons.

- **Chapter 30 Organic Chemistry**

Overview

Organic Chemistry

This chapter introduces the structure and nomenclature of straight chain hydrocarbons, aromatic hydrocarbons, and the functional groups of hydrocarbons.

Organic Chemistry

Outline

The chapter *Organic Chemistry* consists of five lessons that introduce the structure and nomenclature of straight chain hydrocarbons, aromatic hydrocarbons, and the functional groups of hydrocarbons.

- **Lesson 30.1 Carbon, A Unique Element**
- **Lesson 30.2 Hydrocarbons**
- **Lesson 30.3 Aromatics**
- **Lesson 30.4 Functional Groups**
- **Lesson 30.5 Biochemical Molecules**

Overview

In these lessons, students will explore:

- Properties of carbon.
- The definition, naming, and drawing of alkanes, alkenes, and alykynes.
- Compounds that contain benzene.
- Categories of organic compounds that have distinguishing functional groups.
- Categories of biochemical molecules.

Science Background Information

This information is provided for teachers who are just beginning to instruct in this subject area.

Trans Fats

Twenty-first century Americans are becoming increasingly cognizant of the role of dietary fats in their long-term health considerations. Many consumers seek to limit the amount of fats in their diets to minimize their risk of developing coronary heart disease. In particular, one category of fats appears to be linked to several contributory mortality risks: Trans Fats. The chemical structure of this class of compounds consists of long hydrocarbon chains, with one or more trans-configured alkene groups within the chain. Naturally occurring animal fats generally consist of fully hydrogenated or saturated fatty acids, lacking alkene $C = C$ bonds.

In the 1960s, health concerns about saturated fats led to the popularity of unsaturated or partially hydrogenated fatty acids. Also the lower cost of these mainly vegetable oils increased their adoption. These naturally occurring unsaturated fats are usually liquids, due to their predominantly cis-alkene configuration, which produces a "bent" structure that does not pack well. Consumer demand for solid fats, for example, "spreadable" margarines, lead to the application of chemical hydrogenation. Adding hydrogen atoms across the cis $C = C$ bonds can yield the completely saturated fat, or it can also lead to partial hydrogenation. In this case, the $C = C$ bonds do not become fully saturated, but instead the alkene bonds twist to the trans configuration. Unlike the cis-alkene fats, with their bent shape, the trans fats are more linear in shape and like their fully saturated congeners, they pack more effectively and can be produced as solids. Trans fats are also less likely to be attacked by atmospheric radicals and are therefore less vulnerable to rancidity. Their shelf life increases dramatically.

The link between trans fat consumption and heart disease is strongly supported by many medical studies. Other health effects linked to the use of trans fats include liver dysfunctions; due to the synthetic nature of trans fats, they may not metabolize in the same way as other fats.

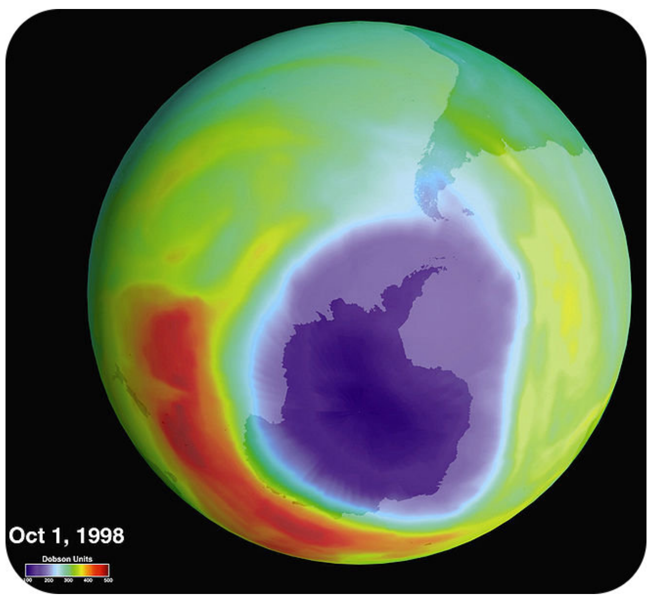
Recently, several municipalities, such as New York City, Philadelphia and San Francisco, have limited or banned outright the use of trans fats in food preparation. The Food and Drug Administration (FDA) now require food manufacturers to list the presence of trans fats on food labels.

Ozone's Role in the Atmosphere

We live on a planet uniquely situated in what astronomers refer to as our solar system's habitable zone. This region can be described as one in which the proper temperature range, elemental composition, and physical conditions have allowed life forms to exist and flourish for billions of years. An appropriate climate range, and the right array of elements, along with sufficient mass for a gravitational, pull enabled Earth to develop a protective atmosphere. The presence of oxygen gas, O_2, causes incoming space debris to burn up usually before reaching the surface. In a similar fashion, an allotrope of oxygen, called ozone, O_3, filters out much of the ultraviolet, B radiation (wavelengths between 280 and 320 nm) reaching Earth from the solar system. This type of radiation is a cause for concern in that it is linked with DNA mutations, particularly those associated with skin melanoma and carcinoma.

$$O_3 + UV \rightarrow O + O_2$$
$$O + O_3 \rightarrow 2\,O_2$$

Nobel Laureate, F. Sherwood Rowland, and his research team at the University of California Irvine, in the 1970s discovered that the amount of ozone found in the stratosphere was diminishing. They found that this depletion could be associated with the cumulative amount of chlorofluorocarbon gases in the atmosphere. These CFC's or as they are otherwise known, Freons, were non-degradable remains of consumer products such as aerosol propellants and refrigerants. Their studies indicated that in the upper atmosphere, ozone was being split in the presence of ultraviolet radiation into diatomic oxygen and oxygen radical atoms, which would, in turn react with the chlorofluorocarbons. The resulting chloride monoxide, ClO, radicals were identified in the upper stratosphere over Antarctica. This location is significant because winter in the Southern Hemisphere (September & October) under the Antarctic vortex, results in the coldest winter temperatures on the planet. Under these conditions, any moisture in that locale exists only in the ice phase. The researchers found that the chlorofluorocarbon/ozone reaction was catalyzed by the presence of certain types of ice crystals in the lower atmosphere. Atmospheric measurements confirmed the accelerated depletion of ozone in the late months of the Antarctic winter, and the increased production of chloride monoxide radicals.



The purple area over Antarctica is low in ozone. The red area is higher in ozone.

Increased ozone depletion has been linked, as mentioned earlier, with an increased risk of DNA mutations. This association has been supported by increasing levels of skin cancer in humans, animals, and even plants, in latitudes where significant ozone depletion is noted. In addition, other conditions, such as diminished immune response and a higher risk of earlier cataract development, appear to be linked with ozone depletion.

Chlorofluorocarbon use has been largely phased out in developed countries, but their relative inertness leads to the situation that their influence will continue to be noted in the stratosphere for the indefinite future. Many manufacturers now substitute HCFC's, (where one or more chlorine atoms have been replaced by hydrogen atoms). These molecules retain many of the desirable properties of chlorofluorocarbons without their contributions to atmospheric destruction.

Pacing the Lessons

Use the table below as a guide for the time required to teach the lessons of *Organic Chemistry.*

TABLE 30.1: 60 Minute Class Periods per Lesson

Lesson	Number of Class Periods
30.1 *Carbon, A Unique Element*	1.0
30.2 *Hydrocarbons*	2.0
30.3 *Aromatics*	1.5
30.4 *Functional Groups*	2.0
30.5 *Biochemical Molecules*	2.0

Managing Materials

The following items are needed to teach the strategies and activities described in the Teachers Edition of the FlexBook for *Organic Chemistry.*

TABLE 30.2: Organic Chemistry Materials List

Lesson	Strategy or Activity	Materials Needed
30.1		
30.2		
30.3		
30.4	Evaluation Activity	Index Cards
30.5		

Multimedia Resources

You may find this additional web based resource helpful when teaching *Organic Chemistry.*

- Introduction to Organic Chemistry: http://www.visionlearning.com/library/module_viewer.php?mid=60

Possible Misconceptions

Making the FlexBook Flexible

An important advantage of the FlexBook is the ability it gives you, the teacher, to select the chapters and lessons that you think are most important for your own classes. You should also consult the standards correlation table that follows when selecting chapters and lessons to include in the FlexBook for your classes.

TABLE 30.3: Standard Addressed by the Lessons in Organic Chemistry

Lesson	California Standards	NSES Standards	AAAS Benchmarks
30.1	10b		
30.2	10d		
30.3	10d		
30.4	10e		
30.5	10a, 10c, 10f		

30.1 Carbon, A Unique Element

Key Concepts

In this lesson students explore the properties of carbon.

Lesson Objectives

- Describe the allotropes of carbon.
- Describe the hybridization of carbon.
- Explain how the hybridization of carbon allows for the formation of large number of compounds containing carbon.

Lesson Vocabulary

hybridization The process of combining sublevels to create a new sublevel.

localized electrons Electrons that are stationary (have fixed positions) between the bond.

delocalized electrons Electrons that are free to move between the bond (in multiple bonding).

allotropes Different forms of the same element based on their bonding.

Strategies to Engage

- Explain to students that there are millions of compounds that contain the element carbon. Have students recall what they know about carbon. Facilitate a discussion with students about how carbon is able to join with other atoms in so many different ways.

Strategies to Explore

- Throughout this chapter, give students the opportunity to create three-dimensional models as much as possible. DI **English Language Learners**

Strategies to Extend and Evaluate

- Have students do library research on Percy Julian, an African American research chemist and prepare a written report, PowerPoint or Keynote slideshow, or display.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions

Have students answer the Lesson 30.1 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Carbon is considered to be unique in the periodic table. What property of the carbon atom makes it unique?
2. What is the difference between sp, sp^2, and sp^3 hybridization?
3. Which of the following are allotropes of carbon?
   a. diamond
   b. buckminsterfullerene
   c. graphite
   d. all of the above
4. The type of bonding in the carbon bonds of ethane (below) are:
   a. sp
   b. sp^2
   c. sp^3
   d. dsp^3
5. The type of bonding in the first two carbon bonds of 1−propene (below) are:
   a. sp
   b. sp^2
   c. sp^3
   d. dsp^3
6. The type of bonding in the first two carbon bonds of 1−propyne (below) are:

$$H-C\equiv C-CH_3$$

   a. sp
   b. sp^2
   c. sp^3
   d. dsp^3
7. What characteristic is responsible for the lubrication property of graphite and buckyball?
   a. localized electrons
   b. delocalized electrons

c. large molecular weight
d. carbon bonding

30.2 Hydrocarbons

Key Concepts

In this lesson students explore the definition, naming, and drawing of alkanes, alkenes, and alykynes.

Lesson Objectives

- Define alkanes as well as name and draw alkanes.
- Define alkenes as well as name and draw alkenes.
- Define alkynes as well as name and draw alkynes.
- Define structural formula.
- Define isomers and be able to draw isomers for alkanes, alkenes, and alkynes.
- Define substituted halogens as well as name and draw substituted halogens.

Lesson Vocabulary

alkanes Compounds containing carbon and hydrogen where the carbon bonds are all involved in single bonding.

saturated compound Organic compound containing all single bonds.

structural formula The formula showing how the bonded atoms are arranged in the molecule.

structural isomers Molecules that have the same molecular formula but different structures.

alkenes Organic compounds containing hydrogen and carbon but contain at least one double bonded carbon atom.

unsaturated compound Organic compound that contain multiple bonding.

alkynes Organic compounds containing carbon and hydrogen and at least one triple bond.

substituted halogens organic compounds where one or more of the branches are a halogen.

Strategies to Engage

Strategies to Explore

- Point out to students that after the first four alkanes in Table 30.1, the names begin with a prefix that identifies the number of carbon atoms in the chain. Challenge interested students to come up with a pneumonic device such as a silly sentence to memorize the names of the first four alkanes.
- Have students set up a three column table to organize the information about alkanes, alkenes, and alkynes explored in this lesson.

Strategies to Extend and Evaluate

- Have groups of students do library research on the future of plastics. Ask some of the groups to find out about plastic superconductors, building materials, or other areas of interest. Have each group give an oral presentation of their findings.

Lesson Worksheets

Copy and distribute the lesson worksheet in the *CK-12 Chemistry Workbook* titled **Organic Nomenclature**. Ask students to complete the worksheet alone or in pairs as a review of lesson content.

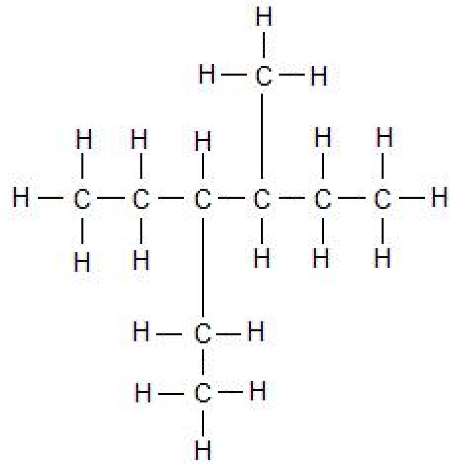
Review Questions

Have students answer the Lesson 30.2 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

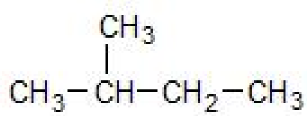
- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Define the terms alkane, alkene and alkyne.
2. What is the difference between a saturated and an unsaturated compound?
3. What is the structural formula of a compound?
4. (1) Which of the following organic compounds is unsaturated?
   a. ethylcyclobutane
   b. 3−ethyl−2−methyl−1−pentene
   c. 2−bromobutane
   d. 2−methyl−1−chlorohexane
5. What is the name of the compound having the following structural formula?



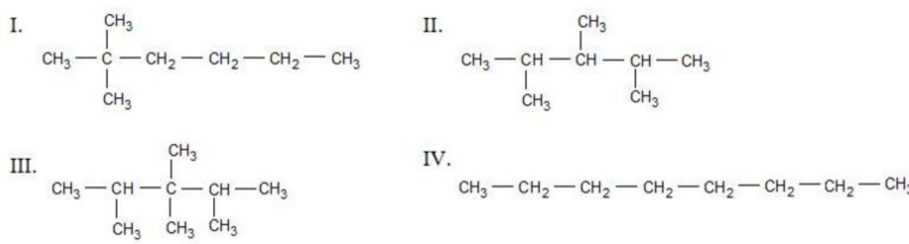
a. 3−ethyl−4−methylhexane
b. 4−methyl−3−ethylhexane
c. 4−ethyl−3−methylhexane
d. 3−methyl−4−ethylhexene

6. Which compound is a structural isomer of the compound shown below?



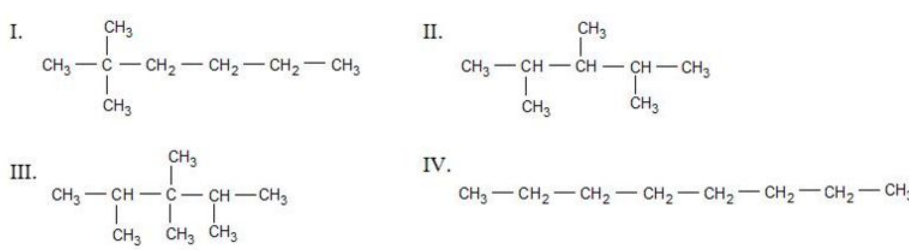
a. butane
b. methane
c. pentane
d. hexane

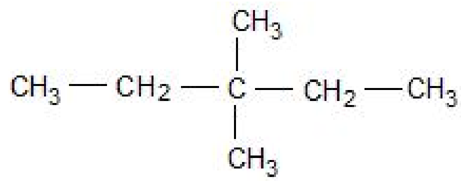
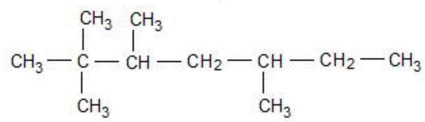
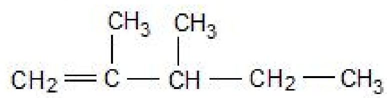
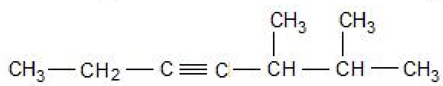
7. Which structures are isomers of the other structures?



a. I, II, III
b. II, III
c. I, II, IV
d. They are all isomers.

8. Which of the following structures has the shortest parent chain?



9. Draw each of the following compounds.
    a. 2,3,4−trimethylpentane
    b. 2−chloro−1−propene
    c. 1−bromo−2−methylbutane
    d. Ethyne
    e. 1−bromo−5,5−dimethylheptane
10. Name each of the following structures.
    a. 
    b. 
    c. 
    d. 
11. Name the isomers for C_6H_{14}.

30.3 Aromatics

Key Concepts

In this lesson students explore compounds that contain benzene.

Lesson Objectives

- Describe the bonding in benzene.
- Define aromaticity.
- Name simple compounds containing benzene.
- Draw simple compounds containing benzene.

Lesson Vocabulary

aromatic A compound contains one or more benzene rings.

benzene ring Equivalent resonance structures representing a 6-carbon ring with alternating $C-C$ double bonds.

hybrid A species with properties in-between the properties of the parents.

resonance To have two of more equivalent Lewis diagrams representing a particular model.

Strategies to Engage

- Have students recall what they know about resonance. Use this opportunity to gauge student understanding, address misconceptions, and generate curiosity for the concepts explored in this lesson.

Strategies to Explore

- Have students create a concept map relating the terms and objectives of the concepts explored in this lesson.

Strategies to Extend and Evaluate

- Have students work in small groups to create an advertisement for a compound containing benzene, such as glue, paint, or furniture solvents. It should resemble an ad that might appear in a newspaper or a magazine. Students should illustrate their ad and write a slogan.
- Have students write a one-paragraph summary of this lesson. Instruct students to correctly use each vocabulary term at least one time in the summary.

Lesson Worksheets

There are no worksheets for this lesson.

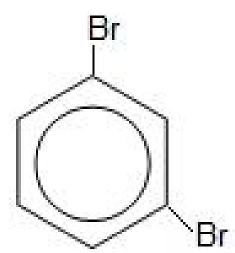
Review Questions

Have students answer the Lesson 30.3 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Define aromaticity.
2. When is the benzene ring referred to as a phenyl group?
3. Name the following structure.



   a. dibromocyclohexene
   b. 1,3−dibromocyclohexatriene
   c. 1,3−dibromobenzene
   d. 2,4−dibromobenzene

4. Which formula represents an aromatic compound?

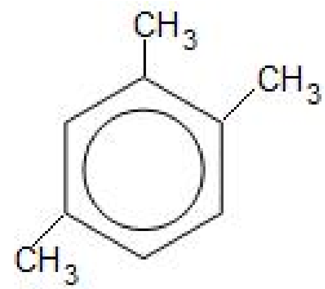
   a. C_2H_2
   b. C_6H_6
   c. C_6H_8
   d. C_6H_{14}

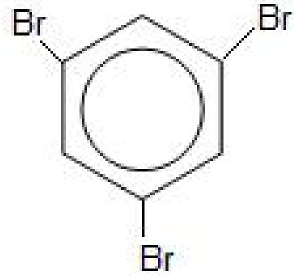
5. How many different possible structures of trichlorobenzene exist?

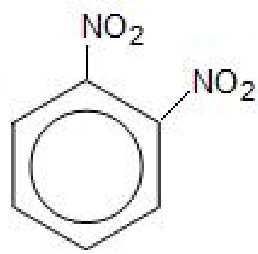
   a. 1
   b. 2
   c. 3

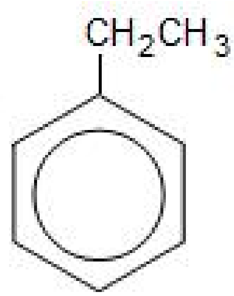
d. 4

6. Name the following structures.

a. 

b. 

c. 

d. 

7. Draw the following structures.

a. fluorobenzene
b. $p-$diethylbenzene
c. $3-$phenylhexane
d. $2-$methyl$-1,4-$diethylbenzene

30.4 Functional Groups

Key Concepts

In this lesson students explore categories of organic compounds that have distinguishing functional groups.

Lesson Objectives

- Identify alcohols, aldehydes, ketones, ethers, organic acids, and esters based on their functional groups.
- Name and draw simple alcohols, aldehydes, ketones, ethers, organic acids, and esters.

Lesson Vocabulary

Alcohol an organic compound in which the hydroxyl group is a substituent on a hydrocarbon.

Aldehyde an organic compound containing the carbonyl group bonded to the end of a hydrocarbon chain.

Ketone an organic compound containing the carbonyl group bonded to a non-terminal carbon atom in a hydrocarbon chain.

Organic acid a hydrocarbon chain ending in a carboxyl group.

Ester an organic compound produced by the reaction between a carboxylic acid and an alcohol.

Strategies to Engage

Strategies to Explore

- Point out to students that while the *-OH* bonding in bases is ionic, the *C-O-H* bonding in alcohols is covalent.
- This lesson includes a description of categories of organic compounds that have distinguishing functional groups. Before reading, prepare less proficient readers by having students write the following on the top of separate sheets of notebook paper:

Alcohols

Aldehydes

Ketones

Ethers

Organic acids

Esters

As they read each section have them write key points under each heading. This will give the students a quick reference and help them to organize the information. Instruct students to write a one-paragraph summary of the information they have read in each section. DI **Less Proficient Readers**

Strategies to Extend and Evaluate

- Write the names of the six categories of organic compounds explored in this lesson on separate index cards. Write their formulas on separate index cards. Have pairs of students compete with each other to correctly match the names with the formulas in the shortest amount of time.
- Have students complete the lab ***Synthesis of Esters***. This lab is located in the Supplemental Lab Book.

Lesson Worksheets

Have students continue with the Organic Nomenclature worksheet started in lesson 30.2

Review Questions

Have students answer the Lesson 30.4 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

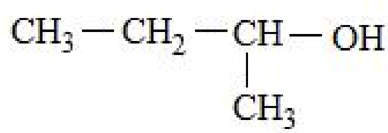
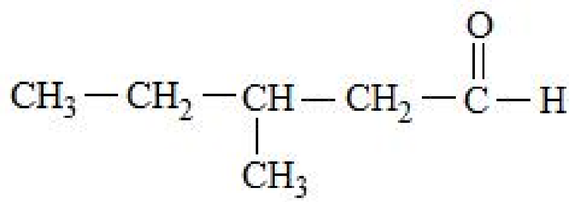
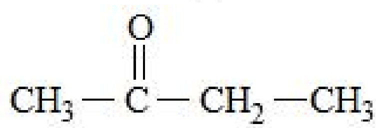
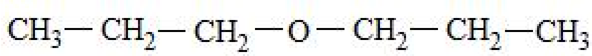
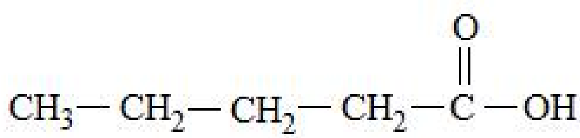
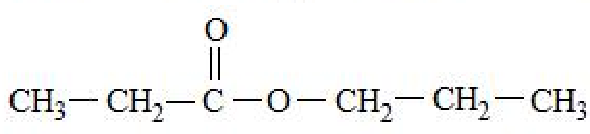
Problem | question=Complete the following chart. |difficulty=Intermediate | solution=

TABLE 30.4:

Group	Distinguishing Feature	Draw Example (with name)
Alcohol		
Aldehyde		
Ketone		
Ether		
Organic Acid		
Ester		

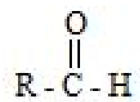
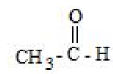
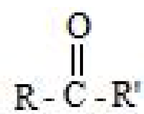
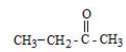
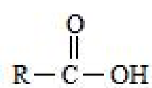
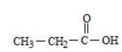
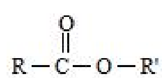
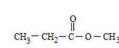
2. What is the difference between the carbonyl group in the aldehydes and the carbonyl group in the ketones? Give an example to illustrate your answer.
3. Which of the following compounds is an alcohol?
    a. CH_3COOCH_3
    b. CH_3CH_2OH
    c. CH_3COOH

   d. CH_3COCH_3

4. To which family of organic compounds does $CH_3COCH_2CH_3$ belong?
   a. alcohol
   b. aldehyde
   c. ketone
   d. carboxylic acid
5. Which class of organic compounds contains a carbon-oxygen double bond? I. Alcohols II. Aldehydes III. Ketones IV. Ethers V. Organic acids VI. Esters
   a. I, III, IV only
   b. II, IV, VI only
   c. I, III, IV, V only
   d. II, III, V, VI only
6. What is the name of the compound represented below?
   a. heptanone
   b. $2,3-$dimethyl $-3-$pentanone
   c. $2,3-$dimethylpentanone
   d. Diisopropyl ketone
7. Name the following compounds
   a. 
   b. 
   c. 
   d. 
   e. 
   f. 
8. (3) Draw the following compounds.
   a. $3-$ethyl [U+0080][U+0093]$2-$ hexanol
   b. $2,2-$dimethylpropanal
   c. $2-$propanone
   d. Dibutyl ether
   e. methanoic acid
   f. methyl butanoate

Review Answers

TABLE 30.5:

Group	Distinguishing feature	Draw Example (with name)
Alcohol	R – OH	$CH_3 - CH_2 - OH$ Ethanol
Aldehyde		
Ketone		
Ether		
Organic Acid		
Ester		

1. The difference between the aldehydes and the ketones is that with aldehydes, the carbonyl group must have one hydrogen atom attached to the carbon atom, with ketones; both groups attached to the carbon atom are alkyl groups.

2. (b)

3. (c)

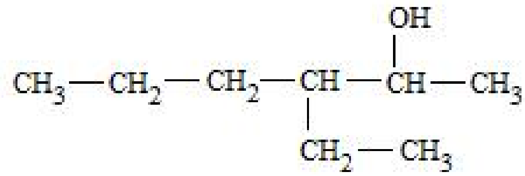
4. (d)

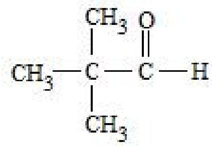
5. (b)

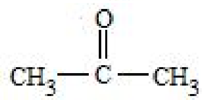
6.

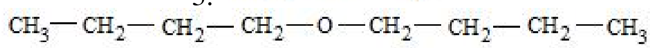
1. 1-methyl -1-propanol
2. 3-methyl pentanal
3. 2-butanone
4. Dipropyl ether
5. Pentanoic acid
6. propyl propanoate

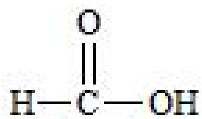
7.

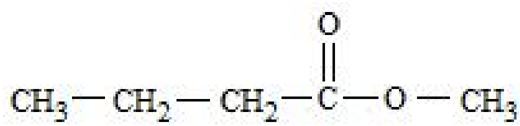
1. 

2. 

3. 

4. 

5. 

6. 

30.5 Biochemical Molecules

Key Concepts

In this lesson students explore categories of biochemical molecules.

Lesson Objectives

- The student will describe the basic structure of fatty acids, monosaccharides, and proteins.
- The student will identify the chemical purpose fulfilled by lipids, carbohydrates, and enzymes.
- The student will describe the biological function of hemoglobin and DNA.

Lesson Vocabulary

carbohydrates Molecules that contain carbon, hydrogen, and oxygen and have the general formula $C_x(H_2O)_y$.

monosaccharide A carbohydrate that is single sugar unit (i.e. glucose).

disaccharide A carbohydrate that is two sugar units joined together (i.e. sucrose).

polysaccharide A carbohydrate that is more than two sugar units joined together (i.e. starch).

lipids Fats and oils (triglycerides) produced for the purpose of storing energy.

fatty acid A carboxylic acid having anywhere from four (4) carbon atoms to 36 carbon atoms.

steroids Compounds where four carbon rings are bonded together with branches and functional groups bonded to the rings.

phospholipids A combination of fatty acids, glycerol and a phosphate group joined together.

polymer A large organic molecule that contains hundreds or even thousands of atoms.

amino acids Molecules that contain an amine group $(-NH_2)$ and a carboxyl group $(-COOH)$.

dipeptide Two amino acids joined together.

polypeptide Many amino acids combined together.

proteins Polymers that are amino acids.

enzymes A subset of proteins that function to speed up a chemical reaction.

DNA (deoxyribonucleic acid) DNA is a polynucleotide that carries our genetic coding; its function is to direct the body in the synthesis of proteins.

Strategies to Engage

- Students are likely to have heard about biochemical molecules in advertising and popular media (e.g., low-fat foods, high-protein diets). Call on volunteers to share with the class anything they already know about biochemical molecules. Point out correct responses and clear up any misconceptions. Tell students they will learn more about biochemical compounds in this lesson.

Strategies to Explore

- On the board or chart paper, outline the main concepts of the lesson as a class. Discuss the main concepts as you prepare the outline.
- Perform the ***Cuprammonium Rayon*** *demonstration. This demonstration is located in the Supplemental Lab Book.*

Strategies to Extend and Evaluate

- As a review of the chapter vocabulary, suggest that students make flash cards, with the vocabulary term on one side, and a drawing of what the term means on the other crossword puzzle.
- Have students research and prepare lists of carbohydrates in food, clothing, and shelter. Discuss findings as a class.

Lesson Worksheets

There are no worksheets for this lesson.

Review Questions Answers

Have students answer the Lesson 30.5 Review Questions that are listed at the end of the lesson in their FlexBook.

Review Questions with Sample Answers

- **Sample answers to these questions are available upon request. Please send an email to teachers-requests@ck12.org to request sample answers.**

1. Fill in the following Table.

TABLE 30.6:

Compound	Main Purpose
Carbohydrate	
Lipid	
Protein	
Enzyme	
DNA	

2. For which biochemical molecule do the triglycerides belong?
   a. carbohydrates
   b. lipids
   c. proteins
   d. enzymes
3. A primary structure is most likely part of what biochemical molecular classification?
   a. carbohydrates
   b. lipids
   c. proteins
   d. enzymes
4. This biochemical molecule is considered a subset of a larger group of molecules?
   a. carbohydrates
   b. lipids
   c. proteins
   d. enzymes
5. Starch is a member of what biochemical molecular group?
   a. carbohydrates
   b. lipids
   c. proteins
   d. enzymes
6. The structure for ribose is shown below. What is the difference between this and that in the DNA molecule?